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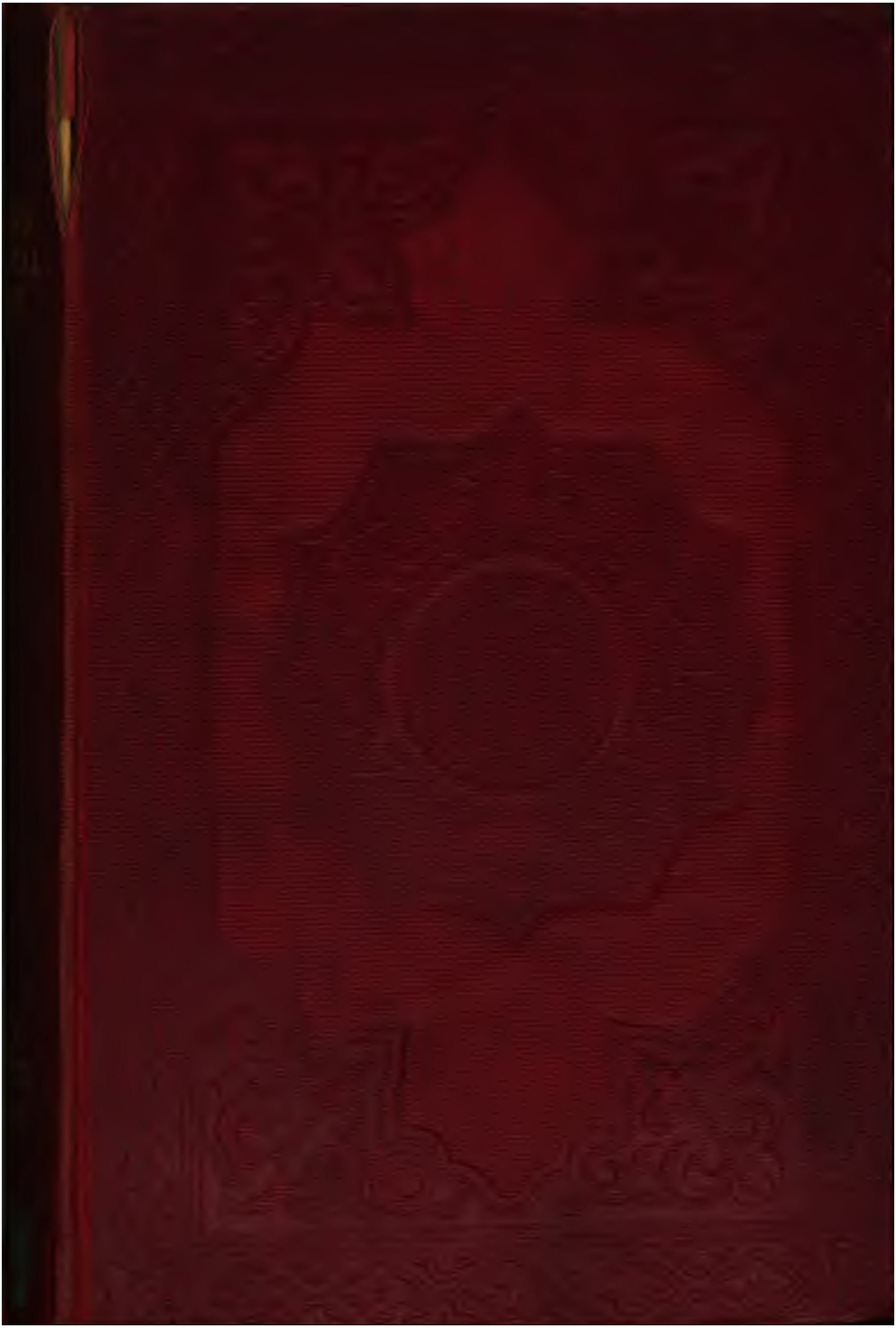

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PREFACE.

"STÖCKHARDT'S Principles of Chemistry," in its English dress, has for many years filled a definite and useful place among elementary text books. It has appealed to the numerous class of students, both men and boys, who in spite of limited means and opportunities are anxious to acquire some experimental knowledge of the science—who intend to *work* at chemistry instead of merely *reading* about it. To such students it is useless to describe experiments which can only be performed with the aid of costly and elaborate apparatus, or with the skill derived from long practice.

The great merit of Dr. Stöckhardt's book is that, while the experiments are clearly described and very numerous, they do not require for their successful performance any but the simplest and cheapest forms of apparatus. It is astonishing what an amount of good work may be done in chemistry with Florence flasks, tumblers, medicine bottles, basins, saucepans, tin plate, iron wire, corks and other articles, which are always at hand. If to these are added some glass and caoutchouc tubing, a few funnels, test tubes and beakers, a mortar and pestle, spirit lamp (if gas cannot be used), a small set of scales and weights, and a measuring glass, the student will be able to perform the great majority of the elementary experiments of chemistry.

Of course other apparatus may be added with great advantage, and apparatus is so cheap in the present day that almost every student will find it possible to make such additions to his stock as he goes on. But it is better not to begin with too much, as the beginner is almost sure to order many things which he will not really require.

When the publishers put Dr. Stöckhardt's work in my hands, I hoped that it might be possible to bring it into accordance with modern chemical ideas, without entirely altering its plan. This soon proved to be impracticable, and I found myself compelled to re-cast, and to a great extent to re-write the book. In doing so I have tried to preserve as far as I could the spirit of the original; but so much has been added, so much subtracted and so much altered, that I have not thought it right to retain the original title of the work. Nearly the whole of the original work which remains consists of experimental details and technical descriptions, which, though they have required much pruning and revision, are as useful now as when first written. A great many new experiments have been added in every part of the work.

The systematic classification adopted in Parts II. and III. has not been attempted in Part IV. (Organic Chemistry), because the more old-fashioned arrangement seemed better adapted for the purposes of a simple experimental study. But I have endeavoured, in two introductory chapters, to give the student some idea of the modern system of classification, and have, in the subsequent pages, referred constantly to these chapters.

In Part I., a somewhat lengthy chapter (Chapter III.) has been devoted to the laws of chemical force. This chapter is longer than it might otherwise have been, because I have thought it right to establish the laws of constitution and chemical action, and the primary meaning of symbols and formulæ, *before* taking advantage of the great assistance of the atomic hypothesis. I trust I have succeeded in proving to the careful student that the symbols and formulæ of chemistry might be used with propriety and advantage, even if the atomic hypothesis were swept away.

In Parts II. and III. I was much assisted by Mr. E. Francis, F.C.S., Demonstrator of Chemistry in the Medical School of this Hospital. It is a pleasure to me to take this opportunity of thanking him for his valuable help in this and in many other things.

C. W. H.

Charing Cross Hospital,
Christmas, 1871.

CHEMICAL APPARATUS.

THE following List comprises most of the apparatus required for performing the experiments described in the present work, and may be purchased by the set of Messrs. JACKSON and TOWNSON, 89, Bishopsgate Street Within, who will also forward more extended lists of apparatus and pure tests, on application, by post.

FIRST SET.—16s.

Packed in a Case, 17s.

1 Holder for Retorts, Tubes, or Flasks.
2 Hard Glass Retorts, 4-oz. and 12-oz. capacity.
2 Hard Glass Receivers, 4-oz. and 12-oz. capacity.
1 Spirit Lamp with single jet, 2-oz. capacity.
2 Hard Glass Flasks for boiling, 4-oz. and 8-oz. capacity.
1 Hard Glass Flask with jet for washing precipitates, 12-oz. capacity, or for burning Hydrogen as generated, or as Marsh's Arsenic Apparatus.
2 Gas Cylinders with ground edges.
2 Ground Glass Plates for the Cylinders.
3 Bottles to hold gases.
1 Hard Glass Tube for preparing Oxygen Gas.
1 Test Tube Stand with 12 holes.
12 Test Tubes, each 5-in. long.
1 Test Tube Cleaner.
2 Analytical Glass Funnels for Filters.
1 Brass Blowpipe.
1 Tripod to support a sand-bath or flask for digestion, or retort for distillation.
2 Wire Triangles for tripod.
2 Iron Dishes for sand-baths, 4-in. and 5½-in. diameter.
3 Porcelain Evaporating Dishes, 2½, 3½, and 4½-in. diameter.
3 Porcelain Crucibles, and Covers, ½ and 1-oz. capacity.
3 Hessian Crucibles, ½, 1, and 2-ozs. capacity.
1 Mortar and Pestle.
3 Cork Borers for perforating Corks.
12 Assorted Corks for Flasks, Tubes, &c.
1 Round File for enlarging holes in Corks.
1 piece of Wire Gauze.
1 ditto of Platina Foil.
1 ditto of Platina Wire.
1 ditto of Copper and Zinc, united, for Galvanic depositions.
1 Iron Spoon for Fusions.
2 Glass Stirring Rods.
3 Tubes for the reduction of Arsenic according to the forms of Berzelius, Clark, and Rose.
2 Bent leading Tubes for fitting up Flasks, &c., for preparing Oxygen, Hydrogen, and other gases.
8 Bent and other Tubes for leading and washing Gases, also as Syphons.
A piece of Wood Charcoal for Blowpipe.

SECOND SET.—25s.

Packed in a Case, 26s. 6d.

Includes the First Set with the extra Apparatus Undermentioned.

1 Stoppered Hard Glass Retort, 4-ounce capacity.
1 Hard Glass Flask, 10-oz. capacity.
1 Pipette or Dropping Tube.
1 Wouiffe's Bottle with three necks, for washing Gases, &c.
1 Spoon with Cap, for Charcoal, Sulphur, &c., when deflagrated in Oxygen.
3 Bohemian Beaker Glasses for Hot Solutions, 4, 6, and 10-oz. capacity.
1 Metal Spirit Lamp with double current and arms to support sand-bath, dishes, crucibles, with Hot Plate and 2 rings.
Where Gas is obtainable a Bunsen's Gas Lamp with Arms and Rings may be substituted for the Spirit Lamp above-mentioned.
1 Thistle-headed Glass Safety Funnel to introduce Acid in the preparation of Gases.
1 Chloride of Calcium Tube for drying Gases.
1 Berzelius' Blowpipe with moveable Platina jet, to prevent Oxidation, and Horn-mouthpiece (*instead of Brass Blowpipe as contained in former set*).
1 pair of Crucible Tongs.
Sheet of 100 Re-agent Labels with Symbols.

THIRD SET.—50s.

Packed in a Case, 52s.

Includes the Second Set with the following additions.

1 Box of Scales and Weights, with glass pans.
1 Graduated 2-oz. Glass Measure.
1 Gas Receiver, 20-oz. capacity, with brass cap, stopcock, bladder-ferrule and jet, for holding or conveying Gases into Bladders, Balloons, &c.
1 Pneumatic Trough, 1-gall.
1 Thermometer, graduated on stem to 400° Fahr., to pass through corks into flasks, &c., *in Wood Case*.
2 Hydrometers for ascertaining the density of fluids, taking 1000 grains of distilled water at 60° Fahr. as standard, one for fluids lighter than water, as spirits, &c., 700° to 1000°, one for heavier fluids, 1000° to 1850°.
1 Solution Tube for holding the fluids for Hydrometer during immersion.
1 Graduated Test Measure, 1000 grains in 100 divisions.

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ARRANGED ACCORDING TO THEIR ATOMICITY.

(See page 84.)

[The names of the Non-metallic Elements are printed in Capitals.]

I. PERISSIADS.—Elements of uneven Atomicity.

1. Always, or generally monad (11).

Name.	Symbol.	Atomic Weight.	Formula as Gas.
HYDROGEN	H.	1·	H_2
CHLORINE	Cl.	35·5	Cl_2
BROMINE	Br.	80·	Br_2
IODINE	I.	127·	I_2
FLUORINE	F.	19·	..
Potassium	K.	39·	..
Sodium	Na.	23·	..
Lithium	Li.	7·	..
Rubidium	Rb.	85·	..
Cæsium	Cæ.	133·	..
Silver	Ag.	108·	..

2. Generally triad or pentad (11).

NITROGEN	N.	14·	N_2
PHOSPHORUS	P.	31·	P_4
BORON	B.	11·	..
Arsenic	As.	75·	As_4
Antimony	Sb.	122·	..
Bismuth	Bi.	210·	..
Vanadium	V.	51·	..
Tantalum	Ta.	182·	..
Niobium	Nb.	94·	..
Gold	Au.	197·	..
Thallium	Tl.	204·	..

II. ARTIADS.—Elements of even Atomicity.

1. Always, or generally, diad (14).

Name.	Symbol.	Atomic Weight.	Formula as Gas.
OXYGEN	O.	16·	O_2 & O_3
SULPHUR	S.	32·	S_2 & S_6
SELENIUM	Se.	79·5	Se_2
TELLURIUM	Te.	129·	Te_2
Barium	Ba.	137·	..
Strontium	Sr.	87·5	..
Calcium	Ca.	40·	..
Magnesium	Mg.	24·	..
Zinc	Zn.	65·	Zn
Cadmium	Cd.	112·	Cd
Indium	In.	74·	..
Copper	Cu.	63·5	..
Mercury	Hg.	200·	Hg
Lead	Pb.	207·	..

2. Often tetrad or hexad (27).

CARBON	C.	12·	..
SILICON	Si.	28·	..
Tin	Sn.	118·	..
Titanium	Ti.	50·	..
Zirconium	Zr.	89·5	..
Thorium	Th.	115·5	..
Aluminium	Al.	27·5	..
Glucinum	Gl.	9·3	..
Yttrium	Y.	61·5	..
Erbium	Er.	112·5	..
Cerium	Ce.	92·	..
Lanthanum	La.	93·	..
Didymium	D.	96·	..
Iron	Fe.	56·	..
Manganese	Mn.	55·	..
Chromium	Cr.	52·	..
Nickel	Ni.	59·	..
Cobalt	Co.	59·	..
Uranium	U.	120·	..
Molybdenum	Mo.	96·	..
Tungsten	W.	184·	..
Platinum	Pt.	197·5	..
Palladium	Pd.	106·	..
Rhodium	Rh.	104·	..
Ruthenium	Ru.	104·	..
Iridium	Ir.	198·	..
Osmium	Os.	199·	..

EXPERIMENTAL CHEMISTRY.

PART I.

GENERAL PRINCIPLES.

CHAPTER I.

MATTER AND FORCE.

PUTTING on one side as unsolved, and probably for ever insoluble, the great metaphysical question whether the various objects and the complex phenomena we observe in nature have any real existence, or are merely images created in our own consciousness; and the farther and scarcely less recondite question whether, even if we admit the external existence of such objects and phenomena, we may not account for them without assuming the existence of any material substance; let us take the ordinary, and what may be called the common-sense view of the universe around us. In doing so, however, we must not venture to assert it dogmatically, or to stigmatise as absurd either of the other views above referred to, startling though they may appear. Both of them have been and are held by many profound thinkers; neither can be proved to be erroneous, neither can be said to be in direct discord with known facts. The thoughtful student will find it good to revert to them from time to time as he advances in knowledge, and as he gains a clearer conception of their meaning he will find them less and less incredible.

Taking then the current view, we hold that all substances known to us, whether in the state of solid, liquid, or gas, consist of *matter* which has an absolute existence. Matter is subject to the operation of certain agents called *forces*, which determine its position in space, and confer upon it its properties. The forces can only be known to us by their operations on matter, just as motion cannot be conceived of except as of something moved. On the other hand, it is no less true that, apart from force, we can form no conception of matter.

Different Forms of Force.—The following are the chief kinds of force.

1. *Motion.*—However motion may have been produced, it is, as long as it remains, a true force. A stone moving through the air is exerting active force, and the planets exert a similar force in their revolution round the sun.

2. *Gravity.*—If a stone is suspended by a string, every one knows that the string is pulled by it. The stone and the earth attract one another, and if the string be cut, the stone will fly to the earth and come to rest. This kind of attraction is called gravitation, and it is said to be due to the force of gravity. It is found to exist in all forms of matter. Solids, liquids, and gases all tend to fly to the earth, and therefore all exhibit the force of gravity, though in very various degrees. Near its surface the great bulk of the earth causes its attraction to be so very much greater than that of any other body that the term gravity is generally applied only to the tendency of bodies to fly to the earth. In other words, gravity is generally taken as identical with *weight.* But it is right to remember that *all* bodies attract one another. One stone attracts another, one drop of water another drop; but the masses of the stones and the drops are so small that the attraction is too slight to be noticed. That the force of gravity keeps the planets in their orbits, and regulates the movements of the heavenly bodies, is known to all. It is the most important doctrine of astronomy.

When the force of gravity is not resisted by some other force equal in power to itself, *motion* is produced.

The special case of gravity that we call *weight* is so important in the study of chemistry that it will be considered separately in the next chapter.

3. *Cohesion.*—The force which binds together the minute particles of which matter consists is called cohesion, but it

is probably identical in nature with gravity. It varies very much in different substances, and even in the same substance in different states. To divide ice into smaller portions requires greater force than to divide water, and no force at all is required to divide steam. In solid bodies cohesion is stronger than in liquids; in gases scarcely a trace of it can be perceived. The *hardness* of a solid is the measure of the force of cohesion among its particles.

When cohesion is exerted between dissimilar substances it is often called *adhesion*. The use of glue is a good example of adhesion.

Experiments.—To prove that the force of cohesion exists in liquids, allow one pan of a pair of scales to rest on the surface of water. A number of weights must now be introduced into the other pan, before the first one will be removed from the water. When at last it rises, some water will still adhere to its under surface. The *adhesion* between the scale-pan and the water is not overcome, but only the *cohesion* between one portion of the water and the rest.

If only half the weights required in the above experiment be introduced, and a little ether be then poured on the surface of the water, the scale-pan will rise at once. The *adhesion* between the ether and the scale-pan is very slight.

Cohesion may also be illustrated by cutting a bullet in half, and pressing together with the hands the two clean and recently cut surfaces. If the cut has been well made, they will unite again with great force. Two pieces of plate-glass will do the same, and the glass-makers have to be very careful not to allow large sheets of plate-glass to lie on one another.

4. *Heat* is an extremely powerful and important form of force. Its influence in chemistry is very extensive, and, like gravity, it will be considered in some detail in the next chapter.

5. *Light* is a force closely allied to heat, and is also closely concerned in many of the changes of chemistry. Almost the whole of the light and heat of our globe comes to us directly or indirectly from the sun.

6. *Electricity.*—*Experiment* 1.—Take a stick of sealing-wax, rub it briskly with a piece of flannel and hold it near some scraps of paper. They will be attracted by it and will fly

upward and adhere to its surface. A tube of glass rubbed with silk will produce the same effect, but the substance called ebonite is more powerful than either. An ebonite comb that has been passed through dry hair will not only attract scraps of paper and small feathers, but in dry weather will emit a crackling sound and a torrent of small sparks, quite visible in a dark room.

In all these cases the strange and interesting force called *Electricity* is set in motion. The "electrical machine" is only a plate or cylinder of glass or ebonite, which, by turning a handle, can be made to rub against a silk cushion covered with amalgam. Electricity produced in this way is called *frictional electricity.*

Another kind of electricity is known, which is called Voltaic, or Galvanic Electricity. The following experiment will illustrate the way in which it is produced.

Experiment 2.—Take a tumbler two-thirds full of water, pour into it gradually an ounce and a half by measure of sulphuric acid (oil of vitriol), and fill the tumbler up with water. The mixture will become very hot, and must be allowed to stand until cold. Very likely a white sediment will form, but this, which is owing to the presence of lead in the acid, will soon subside, and will do no harm. Take a strip of sheet copper, and another of zinc, not too large to go to the bottom of the tumbler, and high enough to reach to the top. Brighten one of the small ends of each with sand-paper, and solder on each bright patch a piece of copper wire about a foot long. It will improve the experiment very much if you can *amalgamate* the zinc plate; that is, coat it with mercury (quicksilver.) To do this the zinc must be cleaned by immersing it for a minute in the weak sulphuric acid, and then, while covered with the acid, pouring a little mercury on it and rubbing it over the surface with a small rag. The surface of the metal will then become as bright as a looking-glass.

The two metals are now to be immersed in the acid with the wires out. They should be about half an inch apart, and must not touch one another. They can be kept in their places by wedges of cork or wood. If the zinc plate has been amalgamated no action will take place *until the ends of the wires are made to touch one another,* when the zinc will immediately begin to dissolve in the acid and electricity to pass along the wires. Such an arrangement constitutes a simple and not very powerful form of *voltaic* or *galvanic* battery. It is called a "simple cell," and the ends of the wires farthest from the plates are called the "poles." If ten such cells are made, and are joined together by thick copper wires, the zinc of one cell to the copper of the next, long wires proceeding from the last zinc at one end and the last copper at the other, a tolerably powerful battery will be obtained, capable of decomposing water and doing other chemical work, but the arrangement is never a very powerful one, and it is better to buy or make one of the more powerful forms of the instrument.

If the two wires from the simple cell are brought to within half an inch of one another, and are then joined with a piece of very fine platinum wire, the platinum will immediately become red-hot. With a more powerful battery a much greater length of wire may be heated to whiteness, and even melted by the electricity. Fine iron wire may be burned in the same way, but not quite so easily.

Magnetism is only a peculiar operation of electricity.

7. *Chemical Force, or Chemical Affinity.*—The last force that we have to consider is called chemical force, because the science of chemistry is almost entirely occupied with its nature and effects. Its operations are spoken of as chemical action.

Every one knows that iron, heated to redness, changes into scales or cinders, and that exposed to moist air or earth, it is converted into rust; that the expressed juice of the grape gradually turns to wine, and this again to vinegar; that wood in a stove, or oil in a lamp, disappears in burning; and that animal and vegetable substances in time putrefy, disintegrate, and finally disappear.

Iron cinders and rust are iron altered in constitution; iron is hard, tenacious, of a greyish-white colour and brilliant; by heating to redness it becomes black, dull and brittle; on exposure to moisture it is converted into a powder of a yellowish-brown colour. Wine is altered *must*, in which nothing of the sweet taste peculiar to the grape-juice can be perceived; but it has acquired a spirituous flavour, together with a heating and intoxicating power, which was not in the *must.* Vinegar is altered wine; it has an acid smell and taste, and has lost its spirituous flavour, as well as its exhilarating properties, its tendency being rather cooling and sedative. Search must be made in the air for the oil and wood which have disappeared during combustion; both these substances are converted into vapour or gas, and heat and light are thereupon evolved with the phenomenon of fire. Of a similar nature are the changes which animal and vegetable substances undergo if kept for a sufficient length of time; they are gradually converted, as they putrefy or decay, into various kinds of gas, some of which emit a very disagreeable odour.

Such processes, by which the weight, form, solidity, colour, taste, smell, and action of the substances become changed, so that new bodies with quite different properties are formed from the old, are called *chemical processes,* or *chemical action.*

Wherever we look upon our earth, chemical action is seen taking place, on the land, in the air, or in the depths of the sea. The hard basalt, the glass-like lava, become gradually soft, their dark colour passes into lighter, they crumble to smaller and smaller pieces, and are finally changed to earth. A potato placed in the earth grows soft, loses its mealy taste, becomes sweet and finally decays. The bud, that sends forth a sickly pale shoot in a dark cellar, when exposed to the light and air grows up a vigorous, firm, and green plant, which, imbibing its nourishment from the moist air and soil, forms from their elements new bodies, not to be found previously in the water or the air. A delicate network of cells and tubes pervades the whole plant, imparting to it firmness; these we call vegetable tissue, or woody fibre. We find in the sap, which passes up and down through these cells, albumin and other viscous substances; in the leaves and in the stalks, a green colouring matter—chlorophyll; and in the ripe tubers, a mealy substance—starch. None of these substances are injurious to health; but if the potatoes grow in the dark and without soil, for instance, in the cellar, there is produced in their long pale shoots a very poisonous body—solanine.

The potato forms one of our most important articles of food. The starch contained in it is not soluble in water, but when received into the stomach, quickly undergoes such a change that it can be dissolved or digested, and then introduced as a liquid into the blood. The blood comes in contact in the lungs with the inhaled air; the blood changes its colour, the air changes its constitution, and the heat which we feel in our bodies is developed. We must conclude, from these changes, that chemical action is going on in our own bodies.

If a piece of iron be heated to redness, till a thick crust of scales is formed around it, and then weighed, it will be found to have increased in weight; consequently, it must have been supplied with something ponderable from the air. This ponderable substance is a species of gas, called oxygen; by its union with the iron it has become fixed, yet by other chemical processes it can be reconverted into its gaseous form. If this crust of iron is now exposed for a time to moist air, it will gradually become rust, and again weigh more than before; it has attracted and united to itself water, and more oxygen from the air. Accordingly, the crust consists of iron and oxygen; the rust, of iron, oxygen, and water, which have become most closely united with each other;—they are chemically combined.

The force which produces such changes as these is called Chemical Force. In the older works on chemistry it is described as "chemical affinity," because it is assumed that the tendency of substances to combine with one another is due to a kind of liking that they have for each other. Iron is so very fond of oxygen that it cannot help combining with it whenever it gets the chance. But it is unsafe in scientific matters to use names which assert more than we really know. Now we know that iron combines with oxygen, but we do not know that there is any "affinity" between the two. We know that force is exerted, but we do not know why. The conditions under which the force is exerted will be discussed in a future chapter.

Relation of the Forces to one another.—The forces we have been considering, unlike as they are in their manifestations, are very closely connected together. Almost any one of them may, under suitable circumstances, be made to produce another, and it has during the last thirty years been shown that the one is directly evolved from the other. It has long been perceived to be impossible, under the present arrangement of the universe, for matter to be created or destroyed. It may change its form and pass ever so many times from one state of combination to another; but the total quantity remains always, as far as our means of observation enable us to judge, the same. The grandest generalization of modern science, a generalization which we may be proud to remember was born within our own day, is that which has taught us that force is equally indestructible; that it also, like matter, may undergo various and countless changes; may appear, now as heat, now as electricity, now as chemical action, and now as visible motion, and yet its sum will always remain fixed and unalterable. There is no new creation of force within the limits of our knowledge. It may lie hid for a time—for any time—but force is force for ever, and whenever we meet with any of its operations in nature we may confidently look for its cause in the disappearance of some previously active form of energy. The birth of one form of force is always the death of some other, and the two are equivalent to one another in quantity.

This grand doctrine, which is known as the doctrine of the conservation of force, requires complex apparatus, and great knowledge for its verification. The demonstration of it is indeed still incomplete, though it has gone far enough to satisfy every rational mind; but a few simple experiments will serve to illustrate the manner in which one force may give rise to others. In other words, we may prove the *correlation* of various forces, though the *conservation* of force must be taken on trust.

Experiment 1.—Rub a piece of copper wire briskly with sand-paper for a minute or two. It will become so hot that the hand cannot bear to touch it, and it will readily ignite phosphorus or the tip of a lucifer-match.

Similar illustrations will occur to all. When the axle of a railway carriage gets dry, so much heat is produced that the carriage is sometimes set on fire. A clever blacksmith can hammer a nail till it is red-hot, and so emits light as well as heat. During the boring of cannon the shavings are too hot to be touched, and, lastly, some savages are able to procure a light by rubbing two dry sticks together.

In all these cases there is a direct conversion of motion into heat, and even, in some cases, into light. Friction is but arrested motion—motion hindered or reduced by an opposing force. If the sand-paper had not pressed on the wire, the same amount of force employed by the arms would have given rise to a much greater motion than that actually produced. The difference, though lost as motion, takes the new form of heat. When a pound weight falls to the earth from a height of 772 feet, its motion is lost, but the weight and the earth receive heat enough to raise the temperature of one pound of water one degree Fahrenheit.

We have already seen that friction may, under certain circumstances, produce electricity, so that we have gained illustrations of the conversion of motion into heat, light, and electricity. Let us now start from heat. The steam-engine affords the best possible illustration of the conversion of heat into *motion*. The work done by the engine is in direct proportion to the heat expended, and so ultimately to the coals burnt. When heat becomes sufficiently intense, it is always accompanied by *light*, and Tyndall has devised a direct experiment in which invisible heat becomes visible light. In fact, heat and light appear only to differ from one another as a low and high tone in music do. The conversion of heat into *electricity* requires for its demonstration a *thermopile*, a somewhat expensive piece of apparatus. It is doubtful whether heat is ever directly converted into *chemical force*, though the converse is one of the commonest of phenomena; but the liberation of the force of *cohesion* during the disappearance of heat can be readily demonstrated, as the following experiment will show.

Experiment 2.—Take a tumbler full of powdered crystals of sodium sulphate (Glauber's salt), and drench it with common hydrochloric acid (muriatic acid—spirit of salt). The salt will rapidly dissolve and become liquid, and the force of cohesion previously hidden or stored-up in the salt will become free and active. But at the same time heat will be taken so rapidly from the materials in the tumbler that the outside will become covered with hoar frost, and if a little water in a test tube be plunged into the mixture it will freeze in a minute or two. Most of the so-called freezing mixtures depend upon this principle.

Experiment 3.—Now try the reverse of this experiment. Take a clean Florence oil flask, nearly fill it with hot water, dissolve as much sodium sulphate in it as the water will dissolve (filtering the solution if it appears dirty) and then, while the solution is nearly boiling, cork the flask tightly and allow it to get quite cold. There is now a great deal more of the solid in the solution than cold water could dissolve, but for some curious reason the excess does not immediately separate out. But take out the cork, and in a minute beautiful feathery crystals appear at the upper surface and pass downwards until the whole mass is solid. If the crystallization does not begin at once, drop in a small crystal of the sulphate, which will immediately produce the desired effect. As soon as the liquid has changed to solid, feel the outside of the flask with your hand. It is sensibly warm, showing that during the exertion and disappearance of the cohesive force heat has been given out.

The conversion of *electricity* into other forms of force has already been illustrated to some extent in the battery. When the platinum becomes red-hot, it emits *heat* and *light*. When the poles of a powerful battery are connected with pieces of coke shaped like lead pencils, and the coke points after touching one another are separated by a short distance, a most intense light, called the "electric light," passes between them. The rays given out by it are similar to those of the sun, and, like the sun's rays, they consist of three kinds:

1. Heat rays.
2. Light rays.
3. Actinic or chemical rays.

The chemical rays are invisible, and are only to be recognised by their effect in producing chemical changes. The interesting art of photography entirely depends on the power of the rays to produce alterations of composition in certain chemical substances, and particularly the compounds of silver. The following experiment will illustrate this curious property.

Experiment 4.—In a room from which daylight is entirely excluded, but which may be lighted by a candle, dissolve sixty grains of silver nitrate (lunar caustic) in an ounce of cold distilled water. Pin a sheet of smooth white paper on to a flat board, and, by means of a flat camel's hair brush, paint it all over uniformly with the colourless silver solution.

Allow it to dry; repeat the painting, and again allow it to dry. Paper so prepared is called "sensitive" paper. To show its properties, place upon it a piece of lace, a few fine feathers, or a sprig of fern, cover it with a sheet of window glass, and press the glass down with weights at the corners. The board may now be brought out of the dark room and exposed to sunlight, or to the strong diffused light of a bright day. The paper will soon begin to change colour, and after a time it will assume a rich brown tint. The whole arrangement must then be taken back to the dark room, the glass, ferns, &c., removed, and the paper immersed in clean rain water and well soaked, the water being changed a good many times. A beautiful image, in white on a dark ground, of the object employed will remain on the paper, which may now be dried and exposed freely to light, for the water will have washed away all the silver salt that has not been affected, and the picture will be permanent.

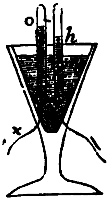
It has already been mentioned that the action of the sun's rays can also be produced by the rays from the electric light, and we have therefore an instance, to which many others might be added, of the production of chemical force from electricity. Even with the single cell described in Experiment 2, page 4, some striking illustrations may be obtained, as will be seen from the experiments which follow.

Experiment 5.—Attach a small strip of platinum foil to the end of each wire. Cut a card so that it will stand upright in a wine-glass, and divide it in two. Then fill the glass with the card in it with a solution of the salt called potassium iodide, and bring one of the platinum *poles* into each of the two divisions which the card makes. The pole which is connected with the copper plate immediately produces a brown colour on that side of the card. The electricity breaks up, or *decomposes* the salt into two substances: potassium, which owing to its action on water is not seen, and iodine, the substance which gives the brown colour in this experiment. Pour into the brown liquid on one side of the card a little of the smooth pasty liquid obtained by adding a good deal of boiling water to starch, stirring and cooling. The starch will combine with the iodine and form a beautiful blue colour.

Experiment 6.—Dissolve in some cold water as much of the salt called copper sulphate (blue vitriol) as the water will take up, and substitute this solution for the potassium iodide used in the last experiment. The card is unnecessary. The salt will be decomposed, and copper will be deposited on the platinum pole which proceeds from the zinc plate. If the poles are reversed the copper will soon dissolve off the one platinum plate and appear on the other. This experiment is a simple example of the important manufacturing process called electroplating. Other metals, such for instance as gold and silver, may by analogous means be deposited from their solutions.

Experiment 7.—*Decomposition of Water by Electricity.*—With the assistance of a more powerful battery, and the little piece of apparatus shown in Fig. 1, water may be decomposed into the two gases of which it consists. Thick platinum wires pass through the sides of a glass and terminate inside in flat plates (stout platinum wires hammered flat at the ends do very well). The holes in the glass can easily be made by a bradawl filed to a square point and constantly moistened with a mixture of turpentine and camphor, and the wires may be cemented in the holes with sealing-wax. The glass is filled with water to which a drop or two of sulphuric acid has been added, and over each of the plates a test tube filled with the same acidulated water is inverted. The two ends of the wires are now connected with the poles of the battery. In an instant bubbles of gas rise from the two plates into the test tubes. The gas which comes from the pole connected with the zinc is called *hydrogen*, that from the other *oxygen;* and it will be observed that the volume of the former is twice as great as of the latter. When a little gas has been collected, the mouths of the tubes may be closed with the thumb, the tubes removed, and the gas examined with a taper in the manner described under their respective names.

Fig. 1.



This is the most important instance of the production of chemical force at the expense of electricity that we possess. The electricity disappears during the experiment, but the hydrogen and oxygen which are produced are able, as we shall find hereafter, to exert an enormous force when they once more combine together.

Lastly, let us study the forces which are called into existence during the exertion and consequent disappearance of *chemical force.* We have, in point of fact, just been considering one of the most interesting cases of this kind in the voltaic battery. In the battery, the electricity which is called into being proceeds directly from, and is proportionate in amount to, the chemical force which is destroyed in the cell. The zinc is constantly being dissolved by the sulphuric acid, and the force which is thereby exerted immediately takes the new form of electricity. From this electricity we can, as we have seen, produce heat, light, and finally, chemical force, so that we can ultimately recover the very same form of force that we lost in the battery. One of Faraday's greatest discoveries proved that there is no loss of force throughout these changes, but that, provided we can prevent the force from assuming other forms, the chemical energy produced by the decomposition of the water is exactly equal to the energy lost in the battery.

One or two simple experiments will illustrate the direct production of other forces from the chemical, but it is unnecessary to multiply them, because the whole study of chemistry is full of such illustrations.

Experiment 8.—Take a piece of phosphorus, about the size of a small pea, and about an equal quantity of iodine. Place them side by side on a slate or a piece of tin-plate, and push the phosphorus with a knife till it touches the iodine. Chemical action is instantly exerted, and the mass bursts into flame, producing thereby *light* and *heat.*

Experiment 9.—Warm two tumblers: put into one a few drops of strong ammonia (spirits of hartshorn), and into the other a few drops of hydrochloric acid (spirits of salt). Each will give out a gas, and when the mouths of the tumblers are held together these gases will combine, and a dense white smoke will be produced, which will before long settle down on the inside of the glasses in the form of a white solid powder (sal-ammoniac). The chemical action is in this case accompanied by the exertion of the force of *cohesion.* The lost force takes the form of heat, which, how ever, can only be perceived by the thermometer.

When gunpowder or any other explosive substance is burnt, it is converted into gas; very energetic chemical action takes place, and a great deal of force is consumed. But the force is entirely reproduced, for the cohesion of the solid is destroyed, a great expansion, which is nothing but a motion of the particles, takes place, and heat and light are evolved. Every one knows what powerful work the expansion may be made to do. Heavy shot can be thrown from a gun to an enormous distance, and solid rocks torn in fragments by the burning even of a small weight of the powder.

The result of the researches of the last quarter of a century has been to convince most scientific men that, various as are its manifestations, there is in nature but one force, namely, MOTION : motion of masses, motion of particles, motion in lines, motions of rotation and vibration, and motion in other modes. As regards heat and light this strange doctrine may be said to be already demonstrated, for these remarkable forces have been proved to be merely peculiar modes of motion in the particles of matter. And the close connection of heat with the other forms of force forbids us to doubt that its nature is essentially similar to theirs, and that they also are modes of motion. Science has still a great work before her in the investigation of force, but enough has already been accomplished to enable us to take a far wider view of the kingdom of nature than was possible to our fathers, and to reveal to us something of the unity and simplicity that underlies the marvellous complexity of the universe.

OBJECTS OF CHEMICAL INQUIRY.

It has already been mentioned that the science of chemistry is concerned with the operations of chemical force. There are *four chief points* to which the attention of the chemist is mainly directed in his study of the solid, liquid, and gaseous substances which are met with in nature.

1. *Their composition.*—Take a piece of bone. How is it affected when strongly heated in a furnace? It becomes whiter, lighter, and less solid than before (bone-ashes). But how is it affected when heated in a covered vessel? It becomes lighter, and black (bone-black). If exposed to boiling water, or to steam, how is it affected? It becomes lighter, and remains white; but in the water is dissolved gelatine. How in hydrochloric acid? It becomes trans-

parent; the bone-earth is dissolved, and a gristly mass remains, which, when boiled with water, turns to gelatine. What is the action of fire upon the gelatine? In a covered vessel it is converted into charcoal, in an open one it burns and disappears. These few experiments show that the bone contains gelatine which is combustible, and an earth which is not so; they show at the same time that it is the carbonized gelatine which, in the second experiment, colours the bone-earth black, and makes it bone-black; that this gelatine dissolves in water, but not in hydrochloric acid, &c. Gelatine and bone-earth are called the *proximate constituents* of bone. but by continued chemical processes these can be resolved still further, that is, separated into simpler constituents. In bone-earth are found phosphorus, a metal (calcium), and oxygen; in the gelatine, besides carbon, three other bodies—oxygen, hydrogen, and nitrogen. These bodies can be decomposed no further by any known method of analysis, and are therefore called *simple bodies*, or *elements*. There are now about sixty known elements, and almost every year adds to their number; but this increase is of little importance to chemical science or its applications, for it consists of elements which but very seldom occur. This separating of compound bodies into simple ones is designated by the name of *decomposition*, and the process of ascertaining the composition of any substance is called *analysis*.

When a substance contains two or more elements, held together by chemical force, it is called a *compound*. Compounds are always quite different in properties from their constituents.

When a substance contains two or more other substances (elements or compounds), not held together by chemical force, but present as it were accidentally together, it is called a *mixture*. In a mixture the properties of the separate ingredients are still perceptible.

This distinction between element, compound, and mixture is very important. The following illustrations will assist to fix it in the memory. Water is a *compound* of the *elements* hydrogen and oxygen, which are held together by chemical force. Air is a *mixture* of the elements oxygen and nitrogen, which are not in chemical combination with one another. Gunpowder is a *mixture*, containing the *elements* carbon and sulphur and the *compound* saltpetre. The *compound* saltpetre contains the *elements* potassium, nitrogen, and oxygen.

2. *The changes which they undergo by the action of other bodies and of the various forces upon them.*—Phosphorus, which is obtained from bones, is luminous in the air, and is gradually converted into an acid liquid; it unites with the oxygen of the air and with water, as the iron did when exposed to moist air. If the phosphorus is gently heated, this union is attended with a vivid combustion, and there is formed a body which is different from the former; to which, if water and lime be added, a new body is formed, very similar to bone-ashes; it is, in fact, artificial bone-ashes. The number of new bodies which may be produced by the union of the elements with each other, or with compound bodies, is infinite; and entirely different substances are often formed, according as the combination takes place under the influence of cold or heat, in water or in air, in greater or smaller quantities. This is *combination* or *synthesis.*

3. *The causes of chemical changes and the laws according to which they take place.*—If chemical experiments are performed as they should be, with the balance in the hand, it will soon be observed, that when two different bodies which can unite with each other are brought together, sometimes a part of the one, sometimes a part of the other, remains free. Further experiments will show how much of one body, in weight, can be united with the other. If all bodies are tested in the same manner, the certainty is finally attained, that all chemical combinations take place only in fixed, unchangeable proportions, and that to every individual body is assigned a definite weight, with which it always enters into any combination whatever. (Chap. III.) This certainty is called a *natural law.* Many such laws of nature have already been ascertained, and they serve as a certain guide to the chemist in his labours, since they cannot, like human laws, be arbitrarily evaded or changed. By them alone we attain to a scientific insight into chemical processes, and to the capability of putting direct questions to bodies by experiment, and of testing the truth of the answers received. An explanation of chemical processes based on natural laws, which presents a clear idea of the subject to the mind, is called a *Theory.*

4. *The extent to which the facts which have been discovered*

may be made useful to man.—When the chemist discovers a new body, or a new property in one already known, or a new method of synthesis or analysis, he imparts his discovery to the apothecary, the physician, the farmer, the manufacturer, and the tradesman, that experiments may be instituted for the purpose of ascertaining whether any advantage, facility, or improvement can be derived for pharmacy, medicine, agriculture, or the arts. Phosphorus ignites spontaneously at a gentle heat; it is used in lucifer matches. Taken into the stomach, it acts as a violent poison; it is at present the most common means for the destruction of rats and mice. The constituents of bone-earth and those of gelatine have been found to be universally present in the seeds of different kinds of corn; the chemist concludes from this that pulverised bones should yield an excellent manure for corn; the agriculturist demonstrates this by experiments on a large scale. In bone-black the property has been discovered of attracting many substances held in solution in liquids, and of condensing them in itself: on account of this property it is used for making impure water potable; the sugar-refiner employs it to make brown syrup colourless; with it the distiller purifies spirit from fusel oil. This is *applied* or *practical chemistry.*

Nothing is better calculated to excite an interest in chemical knowledge than a consideration of the useful application which can be made of it in every-day life. Chemistry teaches the apothecary how to compound and prepare his medicines; it teaches the physician how to cure maladies by means of these medicines; it not only shows the miner the metals concealed in rocks, but aids him also in smelting and working them. Chemistry, in connection with physics, has been the principal lever by which so many arts and trades have been brought to such a degree of perfection within the last few decades, and by its means we have been supplied with numberless conveniences of life that were not enjoyed by our fathers. It cannot be doubted that the farmer must at once regard chemistry as his indispensable friend, for it is this alone which acquaints him with the constituent parts of his soil, with the proper nutriment of the plants he wishes to cultivate, and with the means whereby he can enhance the fruitfulness of his fields.

The ancient so-called Elements.—The dogma of the ancient philosophers, that there are four elements, earth, air, water, and fire, has long been known to be erroneous, although the words are still repeated very often. We now know that three of them are either mixtures or compounds of the true elements, and that fire is but a peculiar operation of heat and light.

CHAPTER II.

GRAVITY, HEAT, PRESSURE.

GRAVITY.

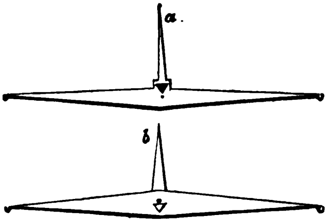
THE force of gravity is chiefly interesting to chemists because it confers upon all kinds of matter the property called weight.

WEIGHING AND MEASURING.

Weighing.—The balance is to the chemist what the compass is to the mariner. The ocean was indeed navigated before the discovery of the compass; but not till after this could the sailor steer with confidence to a certain place, and recover his proper course, however often lost. And so, in chemistry, no systematic method of study could be pursued before the introduction of the balance. The balance is the plumb-line, as well as the touchstone in chemical experiments; it teaches us how to ascertain the true composition of bodies, and shows us whether the questions put, the answers received, or the conclusions drawn from them, are correct or false. Hence it cannot be too strongly recommended to those commencing the study of chemistry to use the balance even in simple experiments. For the experiments described in this book, a common apothecaries' balance is all that is requisite.

Such a balance consists of a brass or steel lever (beam), with arms of equal length, through the centre of which passes a steel wedge-shaped axis, resting on a hardened plate, so that the beam, to the extremities of which the pans are attached, may easily vibrate. It is essential that the axis should be in the right place in the beam, a little *above* its centre of gravity, as in Fig. 2, *a*. The centre of gravity can be found by balancing the beam on its flat side, with the index attached to it, upon a knitting-needle, and when the beam rests horizontally, the point of the needle designates the centre of gravity. If the axis be placed too low *beneath* the centre of gravity, as in Fig. 2, *b*, the beam will overset, if one of the pans is more heavily loaded than the other. If placed directly in the centre of gravity, the balance itself will cease to vibrate when the beam is in an oblique position. When the axis is too high above the centre of gravity, the balance loses much of its sensibility. This latter defect occurs most frequently, but is easily remedied by lowering the axis.

Fig. 2.



Avoirdupois weight is most frequently used in England. The only numbers required in chemistry are the following:—

Pound		Ounces.		Grains.
1	=	16	=	7000
		1	=	437·5

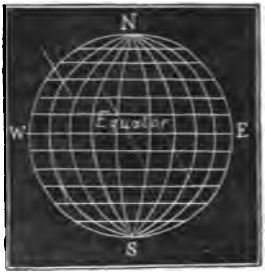
Apothecaries' weight is somewhat different, but the grain is the same.

Pound.		Ounces.		Drachms.		Scruples.		Grains.
1	=	12	=	96	=	288	=	5760
		1	=	8	=	24	=	480
				1	=	3	=	60
						1	=	20

The new French system of weights and measures, which is now almost universally adopted by chemists, is characterised by great simplicity, all its divisions being made by ten; hence the name *decimal* weight and measure. Its unit is derived from the size of our globe.

In order to define the different localities on this globe, imaginary circles, as is well known, have been drawn around it. Those which pass round the earth from east to west, the largest of which is the equator, are called parallels of latitude (circles of latitude); those which pass round the earth lengthwise, intersecting at the poles, *meridians* (circles of longitude). The parallels of latitude gradually become smaller towards the poles; the meridians, on the contrary, are all of equal size. The circle N E S W N represents a meridian or circle of longitude. The fourth part of this circle, or, what is the same thing, the fourth part of the circumference of our earth, as N E, is the basis of the French system. This quadrant was divided into ten million parts, one of which was taken as the unit, under the name of *metre.* A metre is about $39\frac{3}{8}$ inches in length. The smaller measures are produced by dividing by ten, and are designated by Latin prefixes; the larger ones by multiplying by ten, and are designated by Greek prefixes.

Fig. 3.



Smaller Measures.	Larger Measures.
Metre.	Metre.
Decimetre = ·1 or $\frac{1}{10}$ metre.	Decametre = 10 metres.
Centimetre = ·01 or $\frac{1}{100}$ „	Hectometre = 100 „
Millimetre = ·001 or $\frac{1}{1000}$ „	Kilometre = 1,000 „
	Myriametre = 10,000 „

The system of weights was derived from the measure of length, in the following manner. A cubical box was taken, measuring exactly one centimetre in each direction, and this was filled with water at its greatest density (at the temperature 40° F.) (4° C.); the weight of this quantity of water was called a *gramme.* This is taken as the unit of the decimal weights, and is multiplied or divided by ten.

Smaller Weights.	Larger Weights.
Gramme = 1·0	Gramme = 1
Decigramme = ·1 or $\frac{1}{10}$ gramme.	Decagramme = 10 grammes.
Centigramme = ·01 or $\frac{1}{100}$ „	Hectogramme = 100 „
Milligramme = ·001 or $\frac{1}{1000}$ „	Kilogramme = 1,000 „
	Myriagramme = 10,000 „

One gramme is equal to 15·432349 grs. Troy.
One kilogramme is equal to 2·2046 lbs. Av.

It is well enough known that the body whose weight is to be ascertained must be put into one scale, and in the other, weights sufficient to restore the index to its original perpendicular position. The weight of a body thus determined is, in scientific language, called its *absolute weight.* Thus, a piece of sugar weighing two ounces has an absolute weight of two ounces; or, if a vessel be filled with two pounds and one ounce of water, this water has an absolute weight of two pounds and one ounce.

Measuring of Liquids and Gases.—The English system for liquids is as follows:—

Gallon.		Quarts.		Pints.		Fluid Ounces.		Fluid Drachms.		Minims.
C.				O.		fl. ℥.		fl. ʒ.		♏.
1	=	4	=	8	=	160	=	1,280	=	76,800
		1	=	2	=	40	=	320	=	19,200
				1	=	20	=	160	=	9,600
						1	=	8	=	480
								1	=	60

1 gallon of water at 60° Fah. = 277·276 cubic inches, weighs 10 lbs. Av. = 70,000 grains.

1 fluid ounce of water weighs 1 ounce Av. = 437·5 grains.

Grains of water are likewise frequently employed as a measure for liquids. 1,000 grain measures = $\frac{1}{70}$ of a gallon.

Gases are generally measured in cubic inches.

The standard of measure in the French system is the *litre,* which is the same as one cubic decimetre, or a cube of nearly 4 inches on each side. The same prefixes are used for the fractions and multiples as for the metre and gramme, but it will be seen that the litre is also equal to 1000 *cubic centimetres.* Cubic centimetres are generally used in chemistry for the measurement both of liquids and gases. As one cubic centimetre of water weighs one gramme, a litre of water weighs 1 kilogramme. It is almost exactly equal to $1\frac{3}{4}$ pint.

SPECIFIC GRAVITY.

Ice floats in water, iron sinks in it, because the former is lighter, the latter heavier, than water. But if we put a piece of ice in spirit it sinks, or if we put a piece of iron upon mercury (quicksilver), it floats; consequently, ice is heavier than spirit, iron lighter than mercury. It also follows that spirit is lighter than water, since it can support less weight, and mercury heavier than water, as it can bear a greater weight. The terms heavier and lighter, in this sense, correspond to what in scientific language is called *specifically heavier* or *specifically lighter,* and *equal bulks* are always to be understood in speaking of the comparative weights of bodies.

The expression, *ice is lighter than iron*, means, therefore, that, taking equal bulks of each, the former weighs less than the latter; and when we say that mercury is heavier than water, we mean that in equal volumes, as a pint, for instance, the mercury has a greater weight than the water. But in absolute weight, no regard is paid to the volume of subtances.

In order to ascertain *how many times* heavier mercury is than water, or iron than ice, it is only necessary to weigh equal volumes or portions of each, and to compare their weights. If, for example, we take five vessels, each of which would contain exactly 100 grains of water, and fill them respectively with spirit, ice, water, iron, and mercury, the following differences in weight will be found: the vessel with spirit would weigh 80 grains; with ice, 90 grains; with water, 100 grains; with iron (if quite pure), 780 grains; with mercury, 1350 grains.

To facilitate the comparison of the numbers which denote how much greater the specific gravity of one body is than that of another, water has been fixed upon as the standard or unit. Therefore, in the above case, the question is, how much lighter than water are spirit and ice, and how much heavier are iron and mercury? or, in other words, how many times is 100 contained in 80, 90, 780, and 1350? The other numbers, then, are to be divided by 100, the weight of water, and there is found for

Spirit, $\frac{80}{100}$, or, in decimals, 0·80; it is therefore $\frac{1}{5}$ lighter than water.

Ice, $\frac{90}{100}$, or, in decimals, 0·90; it is therefore $\frac{1}{10}$ lighter than water.

Iron, $\frac{780}{100}$, or, in decimals, 7·80; it is therefore $7\frac{8}{10}$ times heavier than water.

Mercury, $\frac{1350}{100}$, or, in decimals, 13·50; it is therefore $13\frac{1}{2}$ times heavier than water.

These numbers represent the *specific weights* (sp. gr.). Thus, according to calculation, spirit having a specific gravity of 0·80, 80 parts of it would occupy the same space as 100 parts of water; therefore it is only four-fifths as heavy as water, or, what is the same thing, one-fifth lighter than water. The specific gravity of mercury being 13·5, that is, $13\frac{1}{2}$ parts of mercury do not take up more space than one part of water; since it is $13\frac{1}{2}$ times heavier than water.

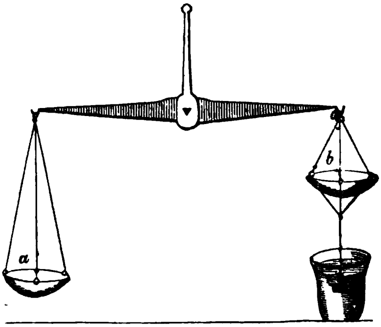
Determination of Specific Gravity.—Experiment.—To determine the specific gravity (the density) of a fluid, a vial is weighed, then filled with water, and again weighed. This gives the weight of the water. Now pour out the water, and refill the vial either with spirit, syrup, lye, beer, or some other liquid, and ascertain by the balance the weight of each. Then divide the weight of each of these fluids by the weight of the water, and the quotient indicates the specific weight. It is very convenient to use a vial made to contain exactly 1000 grains of water, as then, without any calculation, the number of grains which such a vial contains of any liquid expresses its specific weight.

Experiment.—Weigh a bottle filled with water; then place a half-ounce weight (apothecaries' scale) on the pan which holds the weights, and by the side of the bottle nails enough to adjust the beam. Remove both the nails and the bottle from the pan, and put the nails into the bottle. A bulk of water will be displaced equal to that of the nails; to determine its amount, replace the bottle, after it has been thoroughly wiped on the outside, upon the pan, and remove weights from the other pan until the equipoise is restored. The weights taken away (about 32 grains) form the divisor, and the half-ounce, or 240 grains, the dividend; the quotient, $\frac{240}{32} = 7{\cdot}5$, is the specific gravity of iron, of which the nails were made.

The specific gravity of gases is determined by a process which is similar in principle to the above, but is much more difficult. The standard of specific gravity for gases now generally used is hydrogen, which is the lightest gas known. When we say that the specific gravity of oxygen is 16, we mean that it is 16 times heavier than hydrogen. Air is sometimes taken as the standard. It is 14·4 times heavier than hydrogen, so that, knowing the specific gravity of a gas in relation to hydrogen, it is easy to find it in relation to air by dividing by 14·4. Thus the specific gravity of oxygen as compared with air is $\frac{16}{14{\cdot}4} = 1{\cdot}1$.

Experiment.—If we have to determine the specific gravity of a piece of iron, or of any other body which cannot be put into a bottle, it must be fastened by a piece of fine thread to the pan of a common balance (Fig. 4, *b*), the cords of this pan having been previously shortened. Weigh the body first in air, and then in water, immersing it an inch deep. When it is immersed, the opposite pan falls; consequently iron must be lighter in the water than in the air. If the iron in the air weighed half an ounce, then, in order to restore the equilibrium, it will be necessary, as in the former experiment, to remove from the pan *a* 32 grains, equal to the weight of the bulk of water displaced by the iron. The loss of weight is the same, whether the water be removed from the vessel or merely displaced within it. This forms the divisor, and 240, the weight of the iron in the air, the dividend, giving the quotient $\frac{240}{32} = 7{\cdot}5$.

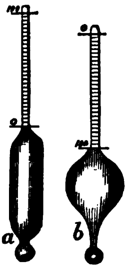
Fig. 4.



Every substance becomes as much lighter in water as the quantity of water displaced weighs; this is a law of nature. If it displaces less water than its weight in the air, it sinks; if more, it floats. Even very heavy bodies can be made to float by increasing their volume; ships are constructed of iron, although it is eight times heavier than water; a tumbler floats upon water, and yet the specific gravity of glass is from three to four times greater than that of water. A thick piece of iron, weighing half an ounce, loses in water nearly one-eighth of its weight; but if it is hammered out into a plate or vessel of such a size that it occupies eight times as much space as before, it then loses its whole weight in water, and will float, sinking just to the brim. If made twice as large, it will displace one ounce of water—consequently twice its own weight; it will then sink to the middle, and can be loaded with half an ounce weight before sinking entirely.

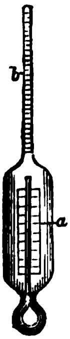
Hydrometer or Areometer.—The same floating body will sink to a greater or less depth in different liquids—deeper in the lighter ones, and not so deep in those which are denser. This has suggested a very convenient instrument for determining the specific gravity of liquids—the *hydrometer*, or *areometer*. This instrument consists of a hollow glass tube, made as represented in Fig. 5. The interior is hollow, and blown out into a bulb at the lower end, to cause it to float; the under part is loaded with quicksilver or shot, to give it a vertical position. The narrow tube serves to denote the depth to which it sinks in any liquid, by means of a scale of degrees, with which it is furnished. There are various instruments of this kind, especially adapted for determining the density of spirits, brandy, oil, lye, syrup, &c. If a hydrometer for weighing spirits is put into water, it sinks only to the lowest point on the scale, 0° (Fig. 5, *a*); but in the strongest alcohol, which is much lighter than water, it sinks to the highest point, 100°. A scale for testing lye (Fig. 5, *b*) must, on the contrary, have the 0° point at the top of the scale, to which it would sink in pure water; for, lye being heavier than water, the instrument would be more or less buoyed up in it, according to its strength. In hydrometers for lighter liquids, the degrees proceed from the bottom to the top; in those for heavier liquids, from the top downwards. In most of these scales, the degrees are arbitrary; and, in order to convert them into the corresponding specific numbers, tables, constructed for the purpose, must be referred to.

Fig. 5.



Experiment.—Pour brandy into a cylindrical jar, and observe the degree which it marks on the hydrometer; then put it in a warm place, and, when lukewarm, again note the degree, which will be higher than before, as the heat has expanded the liquid, made it lighter, and, consequently, apparently stronger than it really is (Chap. III.). The specific gravity of all bodies, when warmed, is less than when cold. On this account, in determining the density of bodies, regard should be paid to their temperature, and it has been agreed to consider 60° F. (15·5° C.) (Chap. III.) as the mean temperature.

Fig. 6.



In the more accurate hydrometers, the mercury serving as the counterpoise has been ingeniously contrived also to indicate the degree of heat of the liquid, by connecting with it a graduated tube. The small scale, *a* (Fig. 6), denotes the temperature, the long scale, *b*, the density. The small scale is frequently so constructed, that the degrees correspond to those on the long scale, and in order to guard against error it is only necessary to add the degrees below the mean temperature to the density, or to subtract from the density those above.

Gold is nineteen times, and silver ten times, heavier than water; gold alloyed with silver must, therefore, have a less specific weight than pure gold. The specific weight of brass is only = 8. Alcohol and ether are lighter in proportion to their purity and strength, while lye, syrup, the acids, &c., increase in density according to their purity. Hence it is evident how important it is, in many cases, to know the specific gravity of a body in order to judge of its quality.

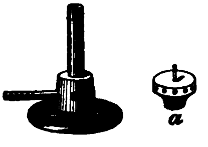
HEAT.

We have already seen that heat is often given out during chemical action. When chemical force is exerted, heat is generally, if not always, produced; and, on the other hand, many chemical changes will not take place without the assistance of heat.

The artificial heat which is employed in chemical operations and in ordinary life is obtained from chemical action. Coal, wood, spirit, oils, and all the other kinds of fuel undergo a chemical change when they burn, which will be fully explained hereafter. They combine with the oxygen of the air, and in doing so chemical force is consumed and changed into heat.

Sources of heat for chemical operations.—If the student is fortunate enough to have gas at his command, he will seldom need any other fuel. The gas may be brought from the nearest burner, from which the jet has been screwed off, by means of an india-rubber tube. For most purposes, the burner known as the Bunsen burner (Fig. 7) is the best. The gas becomes mixed with air in this burner before burning, and burns therefore with a smokeless flame, which gives very little light, but an intense heat. Little iron caps, *a* (Fig. 7), are sold with each burner, which fit on it, and convert the single flame into a ring of smaller jets. A small argand gas-burner with a copper chimney is also very convenient where gentle heat is required.

Fig. 7.

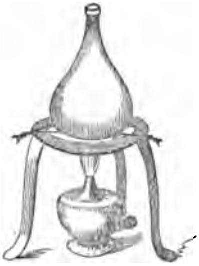


If gas cannot be obtained, a small spirit-lamp, such as shown in Fig. 8, may be used for ordinary purposes, and methylated spirit, which is very cheap, may be burned in it. More powerful spirit-lamps are, however, sold, and small oil-lamps may sometimes be employed for the sake of economy when the heat has to be employed for a long time. Furnaces are scarcely necessary to the beginner in chemistry. The few crucible operations he has to perform may be done in a common fire.

EFFECTS OF HEAT.

Expansion of Liquids.—*Experiment* 1.—Counterpoise a flask—that is, place it on one of the pans of a balance, and equipoise it by weights or shot put into the opposite pan; then fill it with water, and note the weight of the latter. Warm the flask on a tripod over a spirit-lamp or gas-burner, moving it round gently at first, that the flask may heat gradually. The water will soon rise, and part of it run over. When it begins to boil, remove the lamp, and let the vessel cool, and the water will then sink lower than it stood before. How much has been displaced is found by its loss in weight; it will amount to about $\frac{1}{22}$ of the first weight.

Fig. 8.

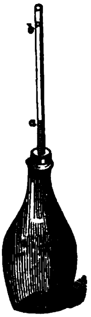


The lamp heats the bottom of the glass vessel, which in its turn communicates heat to the water. The heat expands the water, consequently it occupies a greater space than before, and part of it must run over. Hence it follows that warm water must be lighter than cold water. If a pitcher filled with two pounds of ice-cold water be afterwards filled with boiling water, it will weigh about an ounce and a half less.

The same occurs with all other liquids, and, indeed, also with solids and gases; hence, it may be stated as a natural law, *that all bodies expand by heat, and contract on cooling.* But the amount of expansion is very different in different bodies at the same temperature; alcohol, for example, expands two and a half times more, mercury two and a half times less, than water. When fluids are to be bought and sold by measure, an advantageous application may be made of this principle. If a hundred measures of brandy or alcohol are purchased in hot, and sold in cold weather, there would be a loss of four or five measures; therefore we should gain by buying in winter and selling in summer.

Experiment 2.—In order to observe more accurately the expansion of water by heat, adapt to a flask a cork, rendered so soft by gently striking it with a piece of wood that it may be exactly fitted to the opening by mere pressure; perforate the cork with a round file, and make the hole just large enough to admit a glass tube tightly. Fill the flask with water, so that, when the cork is firmly pushed in, the water shall stand at about *a* (Fig. 9), and heat it as in the former experiment.

Fig. 9.

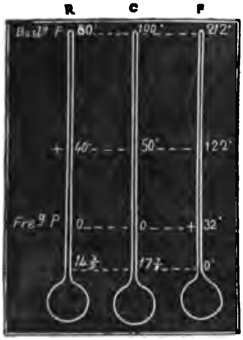


The water, which in the former experiment was displaced from the flask, now rises in the tube, and this higher in proportion to the smallness of its bore. By this means very slight changes of space are rendered visible, and these deviations may be applied to the *measurement* of heat. This is done by particular instruments called thermometers.

Thermometer.—Water might be employed for measuring heat, by marking the boiling and freezing points, and graduating the intervening space; but mercury is far better adapted to the purpose, as it boils and freezes at greater extremes of temperature, and becomes hot and cold more quickly; hence it more quickly denotes the variations of heat and cold: moreover it expands more uniformly than water.

The vessel containing the mercury may also be regarded as consisting of a flask and tube, but which, instead of being joined by a cork, are composed of one entire piece. Having introduced into it a sufficient quantity of mercury, and sealed the open end by fusion, it is immersed in melting ice, and the point to which the quicksilver falls is marked *freezing point;* that to which it rises in boiling water, *boiling point.* The space between these two points can now be divided into degrees, to form the *scale.* The degrees below the freezing point must be of the same dimensions as those above. There are several scales in use, though it is to be regretted that more than one has been adopted. The most common are the three following:—*Reaumur's* (R.), divided into eighty degrees; the centigrade of *Celsius* (C.), into one hundred; and *Fahrenheit's* (F.), into one hundred and eighty degrees. The difference between these can be easily seen in the annexed figure.

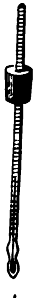
Fig. 10.



According to R. water freezes at 0° and boils at 80°; according to C. it freezes at 0°, and boils at 100°; according to F. it freezes at +32°, and boils at 212°.

Fahrenheit, a philosophical-instrument maker, commenced reckoning very strangely, not at the freezing point, but at 32° below it. His scale is still in common use in England, and the high numbers found in English works are thus accounted for. In Germany, *Reaumur's* thermometer is used, except for scientific purposes, when the *Centigrade*, in common use in France also, is employed. In order to compare these thermometers with each other, it need only be remembered that 4° R. are as large as 5° C. or 9° F. In reducing the degrees of Fahrenheit's scale above 32° to those of the Centigrade, the number of degrees above or below 32° must be multiplied by 5 and the product divided by 9. And to reduce those of Fahrenheit to Reaumur, the number of degrees above or below 32° must be multiplied by 4 and divided by 9. To the degrees above 0°, the sign + is prefixed or understood; to those below, the sign — is invariably added.

Fig. 11.



A cylindrical thermometer, graduated to about 600° F. (315° C.), is best suited for chemical experiments, as shown in the annexed figure, because it can be easily adapted to a perforated cork, and then fitted to a flask, in which liquids are to be heated to a certain temperature. The degrees above the boiling point are to be divided off at distances equal to those below.

Mercury freezes at —38° F. (—39° C.). In the northern regions of the earth a degree of cold of —58° F. (—50° C.)

has been observed, and by artificial means the temperature can be lowered to —148° F. (—100° C.), or even lower. When great degrees of cold are to be measured, alcohol is used in the construction of this instrument, as it does not congeal at —148° F. (—100° C.).

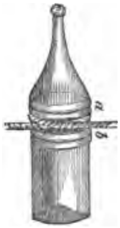
Mercury boils at 662° F. (350° C.); therefore its use must be limited to temperatures below this point. The high temperatures attending ignition are sometimes measured by the expansion of bars of platinum, a metal which does not melt even in the hottest furnace. Such an instrument is called a pyrometer. The determination of very high temperatures is, however, still very unsatisfactory.

Expansion of Solids.—If an iron vessel, when cold, is just large enough to pass through the door of a furnace, it cannot be removed from it when heated. The iron bands or tires of carriage-wheels are applied while red-hot to the frame, and on cooling they contract and bind the wood-work together with great force. A metallic disk, which, when red-hot, fits exactly into a circular box, will, on cooling, become loose, and shake in it. The tire and the disk both become smaller on cooling. These examples show that solids also are expanded by heat and contracted by cold, and explain many of the phenomena of common life. Clocks are apt to go faster in winter, and slower in summer, because the pendulums elongate in summer, and, consequently, vibrate slower; while in winter they become shorter, and vibrate more rapidly. A piano gives a higher tone in a cold than in a warm room, on account of the contraction of the strings. A nail driven into the wall becomes loose after a time, because the iron expands in summer and contracts in winter more than the stone or the wood, and thus the opening is gradually enlarged. For this reason, in the construction of railroads, the rails must not be laid too close together; in the arrangement of steam-pipes, these must not be too firmly inclosed; in roofing, the zinc plates, instead of being nailed *together*, must overlap each other, that they may neither tear nor warp on alternate contraction and expansion.

Brittle bodies, as glass and porcelain, expand or contract so rapidly, by sudden heating or cooling, that they break.

Experiment 3.—Wind round a vial two bands of paper, *a* and *b*, Fig. 12, and secure them firmly with thread; pass a piece of string round the vial, between these folds of paper, and move the vial quickly to and fro on the string until the latter breaks. Then immediately pour cold water upon the place, and the glass will break as evenly as if cut. The sharp edges can be removed with a file. In this manner, common white or green glass bottles may be converted into vessels adapted to chemical and other purposes.

Fig. 12.



It is well known that heat is produced by the friction of two bodies upon each other; that by sliding quickly down a line or a pole by the hands, these will be burnt; and that rapid motion will ignite the axles of a carriage, unless they are well greased. Thus, in the above experiment, the friction produced great heat in the glass, the string emitted a burnt odour and broke, and great expansion of the glass was produced. When the outer surface was suddenly cooled by the cold water, the expanded particles at once contracted, and more rapidly in the external particles than in those of the inner surface, causing the fracture of the glass, and the more easily the greater its thickness. If the temperature had been slowly reduced, it would not have broken.

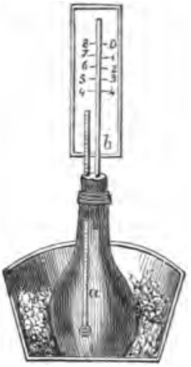
Thus it is obvious (*a*) that glass and porcelain vessels intended for sustaining high temperatures, such as flasks, alembics, retorts, capsules, &c., should be thin, particularly at the bottom; and (*b*) that, when used, they should always be gradually heated and cooled.

The above method of heating glass by a piece of string furnishes the chemist with a simple expedient for removing stoppers which are too firmly fixed in the bottles to be taken out by turning or tapping them. All that is necessary is to wind a piece of thick string round the neck of the bottle, and move it quickly until sufficient heat has been produced to loosen the stopper.

No two solids expand alike; the metals expand the most, and all solids less than fluids.

Expansion by Cold.—A remarkable exception to this law of expansion by heat and contraction by cold occurs in the case of water.

Fig. 13.

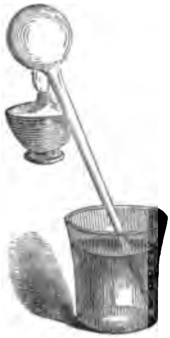


Experiment 4.—A large flask is arranged as directed in Experiment 2, page 29, but inserting a cylindrical thermometer, *a*, through a second hole made in the cork. The flask is filled with water to the top of the tube *b*, and placed in a vessel filled with snow. A strip of paper may be pasted on this tube, upon which the level of the water may be marked as the thermometer falls. The water as it cools will sink in the tube until the mercury stands at 39° F. (4° C.); yet, on cooling still more, it does not fall any further, as we should expect it would, but, on the contrary, it begins to rise again, and continues to do so till it reaches the freezing point. At 32° F. (0° C.) it stands at the same point as when its temperature was at 46° F. (8° C.). Water is accordingly most dense at 39° F. (4° C.); all other liquids continue to increase in density as they cool.

However unimportant this exception may appear at first, our admiration must be the greater when we reflect upon its consequences. Were it not for this, our country would have the climate of Greenland. The freezing of our waters, as the winter sets in, is principally owing to the coldness of the atmosphere. Consequently, the upper part of the water is colder and heavier, and sinks to the bottom; the warmer water ascends, becomes cold, and also sinks. If the water continually became denser, to its freezing point, this circulation would continue till the whole mass of water to its greatest depth reached 32° F. (0° C.), and a few cold days would suffice to convert all our ponds, lakes, and rivers into ice. This does not happen, because the circulation ceases when its temperature has fallen to 39° F. (4° C.), when the water, though yet colder, becomes lighter, and floats on the surface. Thus, freezing can only take place at the surface, and the ice be but gradually formed. At a small depth below the ice the water generally retains the temperature of 39° F. (4° C.)

Expansion of Gases.—*Experiment* 5.—Dip a glass tube, provided with a bulb, into water, and heat the bulb gently; a part of the air is expelled, and escapes in bubbles through the water; consequently, there is not room enough in the bulb for the heated air; but it requires a larger space than it did in its cold condition. It follows from this, also, that the warm air is lighter than cold. If the lamp be removed, the air remaining in the bulb will contract on cooling, and water will be pressed up into the bulb, replacing the air which has been expelled. Such an apparatus, when the tube is small, may be used as an *air thermometer*. When the bulb is heated, the liquid descends in the tube, and *vice versâ*.

Fig. 14.



All gases expand alike when equally heated, differing in this respect from solids and liquids. The extent to which they expand is moreover much greater than with either of the other forms of matter. A gas at freezing point expands $\frac{1}{491}$ of its volume for every degree Fahrenheit ($\frac{1}{273}$ of its volume for every degree centigrade) that is added to its temperature, so that it is very easy, knowing the volume that a certain weight of gas occupies at one temperature, to calculate what its volume would be at some given standard temperature. In experiments on gases their volume is always given (unless otherwise stated) at 0° C. or 32° F., which is therefore said to be the *standard of temperature.*

Example.—A certain weight of gas measures 23 *cubic inches at* 65° *F.: what would be its volume at* 32° *F.?*

491 volumes at 32° become 492 at 33°, 493 at 34°, and so on, increasing 1 volume for every 1° F. Therefore, at 65°, the 491 volumes would become 491 + 33 = 524 volumes. Hence the calculation is a mere rule of three sum:—

Vols. at 65°. Vols. at 32°.
$524 : 491 :: 23 : x = 21{\cdot}5$ cubic inches.

Similar calculations may of course be made in the centigrade scale, as the following example will show.

We have 100 *measures of a gas at* 50° *C.: what would be its volume at* 0° *C.?*

273 vols. of gas at 0° C. = 274 vols. at 1°, 275 at 2°, and

so on. Heated from 0° to 50°, 273 vols. would become 323 vols. Therefore—

$$323 : 273 :: 100 : x = 84 \cdot 5 \text{ measures.}$$

We can in this way find the volume that a gas would occupy at any given temperature, although it most frequently happens that we want to find its volume at freezing point.

We have 100 *volumes of gas at* 17° *C.: what would be its volume at* 193° *C.?*

273 vols. at 0° C. = 290 vols. at 17°, and 466 vols. at 193°. Therefore—

$$290 : 466 :: 100 : x = 160 \cdot 7 \text{ volumes.}$$

These calculations are generally spoken of as "corrections for temperature."

Change of State as produced by Heat.

(1) *From Solid to Liquid.—Melting of Solids.*—Expansion is the first general effect of heat; but in many solid bodies another effect is observed; they change their state of aggregation, they become liquid, they *melt.* Many of them become *soft* before melting, so that they can be kneaded; for instance, butter, glass, and iron; in this condition, glass can be bent and moulded like wax, and iron can be forged.

Experiment 1.—Hold a piece of a small glass tube in the upper part of the flame of a spirit-lamp, revolving it slowly between the fingers; when red-hot, it will be so soft that it can be bent into any shape desired. Thus are easily formed any of the numerous bent tubes required in chemical experiments. For softening larger tubes, a lamp with a double current of air or a Bunsen burner must be used, as this gives a much stronger heat than the simple lamp. To break a glass tube, a scratch is made upon it with a three-cornered file, at the place to be broken, and then it can be parted by gently bending it with both hands.

Most solid bodies become *suddenly fluid,* as ice, lead, &c.

Experiment 2.—Place one vessel containing snow or ice, and another containing a piece of a tallow-candle, on a warm stove, dipping a thermometer from time to time into the melting substances; the temperature will remain stationary in the first vessel at 32° F. (0° C.), in the other at about

100° F. (38° C.), so long as any ice or tallow remains unmelted, but when the melting is complete it will commence rising. The degree of heat at which a body melts is called its *melting point.* Every substance has its own melting point, sometimes above and sometimes below the freezing point; for example, lead melts at 620° F. (327° C.), silver at above 1800° F. (1000° C.), solid mercury at —38° F. (—39° C.). If the two vessels containing the melted ice and the melted tallow are placed in the cold, it will be observed that the tallow soon hardens at about 95° F. (35° C.), but the water does not until the mercury has fallen to 32° F. (0° C.). Thus the congelation of fluids takes place at about the same temperature as that at which they pass from the solid to the fluid state.

Many substances, charcoal for instance, have never yet been melted, and others have never been frozen, as, for instance, alcohol; but it is very probable that, when some method of producing still greater degrees of cold and heat is discovered, we shall succeed in rendering all solid bodies liquid, and all liquids solid.

Latent Heat.—Experiment 3.—Put two vessels of equal size on the top of a warm oven, one of them containing a pound of snow at 32° F. (0° C.), and the other a pound of water at 32° F. (0° C.); when the snow is melted, remove them both. By the touch merely it will be perceived that the snow-water is still cold, while the water in the other vessel has become warm; and the thermometer will indicate that in the former the temperature is at 32° F. (0° C.), in the latter at 174° F. (79° C.). As both vessels have received equal degrees of heat, and when placed upon the oven were of the same temperature; the question suggests itself, What has become of the 142° of heat imparted to the vessel filled with snow? The reply is, This heat has been absorbed by the snow, thus converting it into a fluid,—melting it.

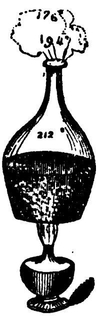
Experiment 4.—Put one pound of snow at 32° F. (0° C.) into the vessel containing the water heated at 174° F. (79° C.), and then examine it with the thermometer; as soon as all the snow has disappeared, the quicksilver will fall to the freezing point. Consequently, the snow has taken from the hot water 142° F. (79° C.) of heat, and has thus become liquid.

The heat which is lost in the melting of ice and other solids is employed in overcoming the cohesion of the solid. The force of cohesion is set free in the manner described in the previous chapter, and when the liquid again becomes solid the heat is reproduced; the exertion of the cohesive force being attended with the evolution of heat. Heat lost during liquefaction is often called *latent heat*, because it was formerly supposed to lie hidden, or storèd up in the liquid. In a limited sense the term may still be retained, for although we now know that heat is nothing but motion, there can be no doubt that the molecules (particles) of the liquid are in incessant motion, which motion they lose to a great extent when the substance becomes solid.

(2) *From Liquid to Gas.—Boiling of Liquids.*—Most liquids, when heated to a certain temperature, different for every liquid, boil and become converted into vapour or gas. It must be understood that there is no real difference between the condition of a gas and of a true vapour, except in regard to the temperature at which they are formed. Many of the so-called gases can by great cold or pressure be converted into liquids; and, on the other hand, when a liquid is converted by heat into vapour, the vapour has all the properties of a true gas. In fact, some substances are known which are liquid in cold weather and gaseous in hot. Gases which become liquid at ordinary temperatures are generally called vapours.

The phenomena of boiling are most conveniently studied in the case of water.

Fig. 15.

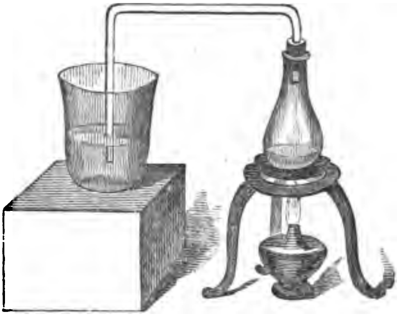


Experiment 1.—Two-thirds fill a flask with spring water, and heat it gently over a lamp (Fig. 15). In a short time numerous little bubbles will appear on the walls of the flask, which will gradually increase in size, and rise towards the surface. These bubbles consist of the gases of air, which are expanded by the heat and expelled from the water. All spring-water contains gases in solution, and to these is chiefly due its refreshing taste, which is not found in boiled water or in that which has been standing for some time. Afterwards, when the water has become quite hot, larger bubbles appear on the hotter part of the flask, which also ascending become smaller, and entirely disappear before reaching the surface of the water; they consist of gaseous water (steam), which condenses as it comes in contact with the cooler liquid above. The collapsing of the particles of water at the places where these steam-bubbles disappear occcasions that peculiar noise which precedes boiling, and which is commonly called the *singing* of the water. When the whole mass of water is heated to 212° F. (100° C.), these bubbles no longer condense, but rise to the surface, where, surrounded by a thin film of water, they remain quiescent for a few seconds, and then, their watery mantle again sinking, they finally burst. This is the *boiling* of water. It boils at 212° F. (100° C.); other liquids boil more readily—alcohol, for instance, at 176° F. (80° C.); others again more difficultly—mercury, for instance, at 662° F. (350° C.).

Latent Heat of Steam.—The space above the boiling water in the interior of the flask appears empty, but it is in fact filled with aeriform water, which has displaced the air that was in it. This aeriform water is called *steam*. It is almost 1700 times lighter than water, because a measure of water yields nearly 1700 measures of steam at 212° F. (100° C.). Within the flask the steam is transparent and invisible, but in the open air it ascends in the form of white clouds, which greatly increase if cold air is blown into the flask by means of a glass tube; for on cooling the transparency of the vapour is disturbed, on account of the formation of drops of water, so small and light as to float in the air. Clouds also consist of this partly-condensed vapour. As the condensation increases, the drops become so large and heavy, that they descend as rain. A thermometer immersed in boiling water indicates 212° F. (100° C.); if placed in the steam immediately above, it shows the same; and this temperature will not rise higher, however long the boiling be continued, or however strongly the flame of the lamp be urged. This is similar to what occurs in the melting of snow; heat disappears, and its disappearance proceeds from the same cause in both cases; steam requires heat for its production, and this heat disappears, or becomes *latent* during its formation, and reappears, or becomes *sensible* during its condensation. We have, in fact, to study the latent heat of steam as well as the latent heat of water.

Experiment 2.—Adapt the shorter limb of a bent glass tube, by means of a perforated cork, to the neck of a flask, and pass the longer limb to the bottom of a beaker-glass or common tumbler. Pour into each of these two vessels five ounces of ice-cold water, and gradually heat the flask upon a tripod until it boils. Note the time required for this operation. Continue the process until the water in the beaker-glass begins to bubble, and note also the time, which will be found the same as that required for boiling the water in the flask. The steam formed in the flask has no other outlet than through the tube into the water, where it condenses, until the contents of the second glass reach the temperature of 212° F. (100° C.), and boil.

Fig. 16.



Both of the vessels must now be weighed; and it will be found that the flask weighs one ounce less, and the beaker-glass one ounce more, than before; consequently, one ounce has passed from the former as steam, and has been condensed again in the latter; and yet this ounce of steam has raised five ounces of water from 32° to 212° F. (0° to 100° C.), that is, through 180° F. (100° C.). Now five ounces raised 180° is the same as one ounce raised 180 × 5 = 900°, so that the ounce of steam in condensing gave out 900° F. (500° C.) of heat. This is expressed by saying that the latent heat of steam is 900° F. (500° C.). More exact experiments give as the true number 967° F. (537° C.). The property of steam to give out a large quantity of heat during condensation peculiarly adapts it for the heating of other bodies, the burning of them being thus guarded against, as the heat of steam in open vessels can never exceed 212° F. (100° C.). Apothecaries avail themselves of steam in the preparation of infusions, decoctions, and extracts; it serves for many of the processes of cookery, and for the distillation of spirits and other liquids; it is employed in dyeing and bleaching establishments, and is often resorted to for heating apartments, buildings, laundries, &c.

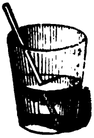
All gases give out heat when they condense into liquids, but the amount of heat so evolved is only known in a few cases. Steam gives more than any other known gas.

Evaporation.—Aqueous Vapour.—Water exposed in a vessel to the open air disappears in summer more rapidly than in winter; the heat of the air converts it into gas—it *evaporates.* The same happens as in evaporation over the fire, only in the former case evaporation takes place without any visible motion of the water, owing to its becoming aeriform, not throughout the whole mass at once, but upon the surface only. The vapour rises in an invisible form in the air. Warm air, indeed, takes up more of it than cold, but a fixed quantity of it only for each temperature. Thus one hundred measures of air at 32° F. (0° C.) absorb two-thirds of a measure of vapour; at 50° F. (10° C.), one measure and a quarter; at 68° F. (20° C.), two and an eighth measures, &c. If the air has not absorbed all the vapour which it can, it eagerly takes up more, as, for example, when one hundred measures of air at 68° F. (20° C.) contain only one or one and a half measures of vapour; it is then called *dry air*, and wet articles are soon dried in it by rapid evaporation. But if it be already saturated with vapour it is called *moist* air; and damp articles cannot be dried in it, or at least but slowly. If yet more vapour be added to this saturated atmosphere, or if it be cooled, then the excess separates in visible particles, called *mist* or fog when they lie near the surface of the earth, and *clouds* when they float in the higher regions of the atmosphere. The white smoke which in winter is seen rising from the chimneys, the visibility of the breath in frosty weather, and the smoking of rivers in winter and after a storm, are phenomena of the same kind.

If the cooling of the air is occasioned by a cold solid body, the vapour is then condensed in small drops of water, as may be observed on the outside of a cold glass when brought into a warm room, and the deposit of moisture on the inside of our window-panes, when cooled by the external cold air. The temperature at which this occurs is called the *dew-point*, signifying the temperature at which the air is saturated with vapour.

Experiment 3.—Fill a tumbler one quarter full with cool water, place in it a thermometer, and at short intervals gradually add ice or cold water, until moisture begins to deposit on the outside of the glass. Then observe the degree indicated by the thermometer, which is the dew-point. If much cold water must be added before the glass clouds over, that is, if the dew-point is much lower than the temperature of the air, fair weather may be expected; while, on the contrary, if the difference between the dew-point and the temperature of the air be but slight, rain may soon be expected, as then the air requires but a slight addition of moisture or increase of cold to become saturated. Instruments by means of which the amount of moisture in the air is ascertained are called *hygrometers*. Many substances readily imbibe moisture from the air, and become damp; such bodies, for instance, as catgut, carbonate of potassium, sulphuric acid, fresh barley-sugar, &c., are called *hygroscopic*.

Fig. 17.

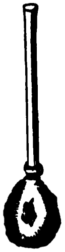


Evaporation may be accelerated, not only by heat, but also by a current of air, because by this means the air above the surface of the fluid, which is charged with vapour, is removed and replaced by a drier, and, as it were, more thirsty air, which takes up the vapour more rapidly and abundantly than the former. For this reason, the earth dries rapidly after rain, when followed by a high wind, and hence it is necessary in kilns, laundries, drying-rooms, &c., to arrange them in such a manner that the air, when saturated with moisture, may be constantly replaced by dry air.

That *heat disappears* during slow as well as rapid evaporation may be readily illustrated by the following experiment.

Experiment 4.—Fill a tube half full of water, and fasten securely round the bulb of it a piece of wadding; saturate the wadding with cold water, and then twirl the tube rapidly between the hands; presently the water in the tube will become sensibly colder, and the degree of cold may be accurately determined by the thermometer. Moisten the wadding with ether, a very volatile liquid, and twirl it again in the same manner as before; by which means its contents, even in summer, may be converted into ice. Water evaporates slowly, ether rapidly; and both require heat for their conversion into vapour, and in the above experiment they obtain this heat from the water in the bulb, which is of course the reason of the water becoming cold. On this principle, one feels cool on just leaving the bath, when invested in damp garments, or when the floor of a hot apartment is sprinkled with water. It explains, also, how man is enabled to support the scorching sun of the hottest climates, and even to endure a heat of 212° F. (100° C.), without his blood exceeding the temperature of from 100° F. (38° C.) to 104° F. (40° C.); it is owing to the more copious perspiration, which, by evaporation, renders all the heat above 104° F. (40° C.) latent. If we blow on hot soup, it is also the increased evaporation which cools it more rapidly; but if we blow on the cold hands in winter, they become moist and warm, because the latent heat contained in the vapour of the breath is set free, as the vapour is condensed into water.

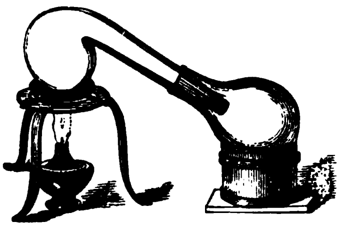
Fig. 18.



Distillation.—If the condensation of watery vapour be carried on in a closed vessel, the water may be collected as it forms.

Experiment 5.—A small glass retort is half filled with water, and heated; the steam, as it forms, passes through the neck of the retort into a glass receiver, contained in a vessel filled with cold water, and is there condensed. That the refrigeration may take place more rapidly, the receiver is covered with coarse blotting-paper which is frequently moistened with cold water. This operation is called *distillation* (from *distillare*, to drop), and the pure water obtained is said to be *distilled*. It is purer than spring-water, for this reason, that the non-volatile, earthy, and saline portions contained in all spring-water do not evaporate, but remain in the retort. By this means also

Fig. 19.

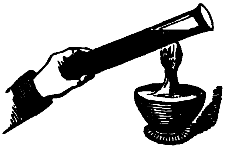


very volatile bodies can easily be separated from less volatile ones; as in the distillation of brandy, where the more volatile spirit is separated from the less volatile water. Copper stills are usually employed for distillation on a large scale, and for condensers vats are constructed, holding serpentine pipes, or *worms*, which present a greater condensing surface than if the pipe had passed straight through the vat. The cold water with which the vats must be filled is very soon warmed by the heat liberated in the condensation of the steam, and must occasionally be renewed by leading off the hot water from above, and letting in a fresh supply of cold water beneath.

DIFFUSION OF HEAT.

Conduction of Heat.—*Experiment* 1.—A test-tube, nearly filled with water, is held over a spirit-lamp, in such a manner as to direct the flame against the upper layers of the water; the water will boil at the top, but remain quite cold below. If mercury is treated in a similar way, its lower layers will gradually become heated. The particles of mercury will communicate the heat to each other, but not so the particles of water. Substances through which, as in mercury, heat rapidly passes, are called *good conductors;* but bodies which comport themselves like water are called *bad conductors* of heat. In the former class are included principally the metals, and in the latter, stone, glass, wood, snow, water, and especially soft bodies, as cloth, fur, linen, straw, paper, ashes, &c., and also the gases.

Fig. 20.

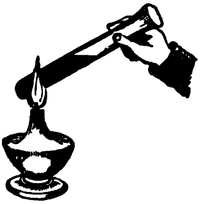


The *good conductors* are readily heated, and soon become cold again, as is well known to be the case with iron stoves. A piece of iron feels hotter in the sun and colder in the shade than a piece of wood at the same temperature. The explanation of this delusion of the sense of touch is, that the warm iron conducts the heat more rapidly to the hand, while the cold iron withdraws it more rapidly than the wood is capable of doing.

The *bad conductors* of heat become only slowly heated, and also slowly cooled; for this reason, stoves constructed of brick (the Russian stove) and those made of Dutch tiles, a preparation of clay, retain their heat longer than iron stoves. Bad conductors are frequently employed both for preventing the quick heating and the quick cooling of bodies. Vessels of glass and porcelain are placed on sand (a sand-bath) or ashes, to heat them gradually, and thus guard against their breaking. If a hot liquid is to be poured into them, it must be done by small portions at a time, twirling the vessels round for some minutes before adding more.

On removing vessels from the fire, the precaution should be taken never to place them while hot on metal or stone, but always on some bad conductor, such as straw (straw rings) wood, paper, cloth, &c.; as they are often cracked by sudden cooling and contraction, which is frequently produced even by a current of cold air. Doors of furnaces, ladles, &c., are provided with wooden handles, to prevent those using them from being burnt. Should a person desire to hold a flask or a test-tube while liquids are boiling in them, he must wrap round them several folds of paper, or tie round them a piece of twine, in order that they may serve as a bad conductor between the glass and his fingers. By inclosing substances in bad conductors, the entrance of cold, or, more correctly, the departure of heat, may be prevented; this principle is illustrated in our method of clothing, in the protection given to our wells and trees by covering them with straw, in the preservation of plants by snow, and in numerous other phenomena of daily occurrence. Hence bad conductors are frequently called preservers of heat.

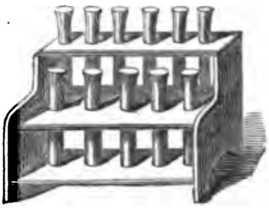
Fig. 21.



Convection of Heat.—Liquids and gases, though very bad conductors, readily become heated by a process called *convection*, which is a necessary consequence of the expansion produced in them by heat.

Experiment 2.—Water, to which some sawdust has been added, is heated in a test-tube over a spirit-lamp. The tube is held by the upper part, and rotated for some minutes between the fingers, that the flame may have equal access to all the lower parts of the tube. If the water be carefully observed, it will be seen that the sawdust ascends on the upper surface of the liquid, and descends in the lower strata; the warm water expands, and, becoming lighter, rises upwards, while the colder, and consequently heavier, water sinks; the water *circulates*. In consequence of this circulation, the heating of fluids takes place more rapidly when the heat is applied *beneath*. If the *upper part* of a test-tube full of water is heated by the lamp, the water will boil at the top while it remains quite cold, and may even contain ice, held down by a wire, at the bottom. Test-tubes are cylindrical glass vessels with rounded bottoms. To prevent their breaking on the application of heat, the bottom must be thin, and blown of a uniform shape. A simple wooden rack, as in the annexed figure, serves as a convenient stand for them.

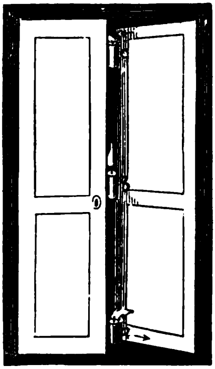
Fig. 22.



Convection in Gases.—Currents of Air.—A great many phenomena of daily occurrence may be explained by the difference in lightness between warm and cold air. When a fire is lighted in a stove for the heating of an apartment, the air immediately in contact with the stove is first heated, becomes lighter, and ascends; colder air rushes in to supply its place, and this likewise becomes heated and ascends; consequently a constant circulation of air is kept up. By a similar circulation, the whole atmosphere of the earth is kept in continual motion. At the equator the strongly heated air ascends and moves in the upper regions of the atmosphere towards the poles, while in the lower regions the current of cold air flows from the arctic zone towards the equator, in order here to restore again the equilibrium, disturbed every moment by the ascent of the warm air. These regular currents of air, the direction of which is somewhat diverted by the revolution of the earth on its axis, are called trade-winds.

In every heated apartment, a difference between the heat of the air near the ceiling and that near the floor is very perceptible. If a door or window in such a room be opened, a current of air is produced, the direction of which may easily be perceived by holding a lighted candle in the opening; the flame, when held above, at *c* (Fig. 23), is blown *from* the room; when placed below at *a*, it is blown *into* it; consequently, the light warm air above rushes out of the room, and is replaced by heavier and colder air from below. A draught of air is also noticed in passing from the sunshine into the shade; where the sun shines, the warmer air ascends, and the colder air from the shade supplies its place. For the same reason, a current of air is produced wherever a fire is burning, in every stove, and round every lamp. To *ventilate* a room properly, the hot and foul air should be allowed to escape from some aperture near the ceiling, while cool and fresh air is admitted near the ground.

Fig. 23.



The air-balloons, first constructed by Montgolfier, strikingly show how buoyant air may be rendered by heat; these are caused to ascend merely by filling them with air, kept continually hot by a fire beneath.

Radiation of Heat.—By conduction, bodies can communicate or abstract heat only when in contact. But heat is felt even at some distance from a fire or from a heated stove, and the earth is warmed by the sun, although a space of millions of miles is between them. This sort of heating is called *radiation of heat.*

Experiment 3.—Envelop three tumblers with paper, one with silver paper, another with white, and a third with dull black paper, and place them in the sun; a thermometer will indicate that the tumbler with the black paper is heated the most, and that with the silver paper the least, and yet all these vessels have been equally exposed to the rays of the sun. This difference is explained on the principle, that some substances have the power of absorbing the rays of heat, while others throw them off again, or *reflect* them. Good reflectors of heat, such as bright metals, are the worst absorbers, for a thing cannot absorb the heat and also reject it.

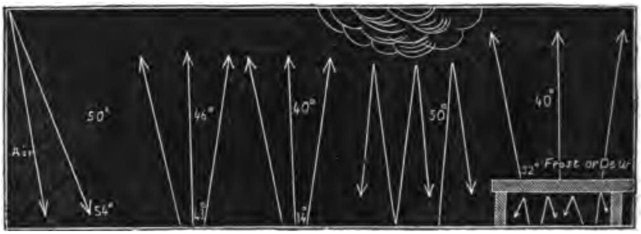
If hot water is poured into the tumblers enveloped with paper, and the cooling of it noted by the thermometer, the contrary effect will be observed; the glass covered with black paper will first become cold, and that wrapped in silver paper the last; because bodies with dull surfaces radiate the heat more rapidly than those with polished surfaces. Good absorbers of heat are therefore seen to be good radiators, and good reflectors bad radiators. For this reason, coffee retains heat longer in a bright than in a tarnished pot; a stove of glazed Dutch tiles remains hot longer than another of unglazed tiles; a smooth sheet-iron stove, longer than a similar one of rough cast-iron, &c.

The radiation of heat enables us to explain some of those common natural phenomena which otherwise would remain obscure. Why do not the rays of the sun, even in the hottest summers, melt the snow upon the tops of high mountains, which are nearer than the level portions of the earth to the sun? Because they only heat those bodies which can absorb their warmth, as the rough surface of the earth. The snow is indeed struck by the rays of the sun, but being a white and shining body it reflects them and remains cold.

Formation of Dew.—When the surface of the earth has become warm, the air is heated by it; hence, during the day the lower strata will always be warmer than the upper. But a change takes place after the sun has gone down. The earth continues to radiate heat without receiving any in exchange, and its temperature consequently diminishes. On the other hand, the air does not so readily part with its heat, and therefore it retains during the night a higher temperature than the surface of the earth; it is only cooled where it comes in contact with the colder earth. If this cooling should reach the dew-point of the air, then the vapours are condensed, on the soil or on the plants growing upon it, in the form of small drops, just as a tumbler is covered with vapour when brought from a cold into a warm room,—dew forms. If the temperature of the earth sinks in the night to the freezing point, or below it, the aqueous vapour is deposited in a solid form, and is called *hoar-frost.*

The radiation of heat from the earth is greatest when the weather is clear and serene, but it is obstructed by clouds and wind. Hence the most copious deposit of dew takes place only in clear and quiet nights. The clouds serve as a screen in reflecting back the rays of heat to the earth, so that it is but slightly cooled. The same effect is produced by the mats, straw, and boards with which the gardener covers his young plants, to protect them from the late frosts of spring, or from freezing. The annexed figure, in which arrows denote the direction of the rays of heat, will serve to render this process more intelligible.

Sunbeams. Fig. 24.



59° Surface of the earth.	41° Dew.	32° Frost.	54° No dew or frost.	41° No dew or frost.
In the day-time.	In clear and serene nights.		Cloudy or windy nights.	Clear nights. Soil protected.

SOLUTION AND CRYSTALLIZATION.

These subjects are so closely connected with that of heat that they may conveniently be studied in this place.

Solution.—Water can dissolve many bodies, and unite intimately with them, without losing its transparency. Such combinations are called solutions. If rain-water meets with soluble substances, either in the earth or in the rocks through which it oozes, it dissolves them; and this explains why almost all spring-water, on evaporation, yields an earthy or saline residue. Frequently this residue, particularly when it contains lime, is so altered during evaporation, that it can no longer be dissolved in water, and forms a hard incrustation round the inner sides of the vessels used in cookery. The springs of Carlsbad deposit so much residue, that articles immersed in them appear in a short time to be externally petrified or incrusted. If water is unusually rich in soluble substances, especially such as possess medicinal properties, as, for example, iron, sulphur, &c., it is called *mineral water*, and the springs from which it issues are called mineral springs. A pound of sea water contains about half an ounce of substances in solution.

Experiment 1.—Pour a teaspoonful of slaked lime into a bottle, and fill it with water; cork it up, and, after shaking it for some minutes, let it stand until the water has become perfectly clear. By carefully inclining the bottle, most of the liquid may be poured off free from the sediment. This operation is called *decantation*, and the clear liquid is *lime-water*. Lime is but slightly soluble in water, seven hundred and seventy-eight ounces of water being required to dissolve one ounce of lime; the excess remains undissolved, and as lime is heavier than water, it settles at the bottom. That a portion of it has been dissolved is known by the peculiar taste imparted to the liquid. This taste is called *alkaline.*

Keep a part of the lime-water in a well-stopped bottle for future use; it will remain transparent and clear. Pour the remainder into a tumbler, and expose it to the air; the water soon becomes turbid, and covered with a film, which gradually grows thicker, and settles at the bottom. When, after some days, the water has become clear again, it will have lost its alkaline taste; the lime dissolved in it, having been chemically changed by the air and rendered insoluble, will be found as a powder at the bottom of the tumbler.

Experiment 2.—Put into a flask half an ounce of litmus; pour over it three ounces of water, and let it remain in a warm place until the liquid has obtained a dark-blue colour. Litmus consists of a blue colouring-matter, which is soluble in water, and is hence taken up by it; it also contains some earthy matter which is insoluble, and is deposited as a slimy mass. These two substances might be separated from each other, as in the former experiment, by decantation, but it can be done more readily by *filtration*. For this purpose cut a piece of blotting-paper into a

Fig. 25.

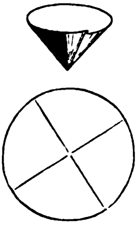
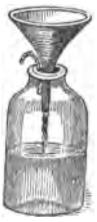


Fig. 26.



circular shape, fold it together twice, and then place this filter into a glass funnel. That the paper and the glass may not come into too close contact, place between them thin pieces of wood or glass; a piece of string must also be inserted between the funnel and the neck of the flask into which the liquid is to be filtered, to allow an opening for the escape of the air from the flask, as otherwise the fluid could not flow in. The filter, which must never be higher than the top of the funnel, is first moistened with water before the fluid is poured upon it. Blotting-paper consists of fine linen or cotton-fibres matted together, between which are small interstices or pores, through which liquids, but no fine solid particles, can pass; these remain on the filter. Writing-paper cannot be used for filtration, as its pores are filled up by size or starch.

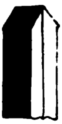
Experiment 3.—Pour a part of the solution obtained into a saucer, and pass strips of fine blotting or of letter-paper one or more times through it, until they have acquired a distinct blue colour. Preserve these strips, after being dried in a box; they are called blue litmus, or test-paper; they are reddened by vinegar, lemon-juice, and all acid fluids, and serve to test a liquid, to ascertain whether it is *acid* (has an acid reaction).

Experiment 4.—Mix cautiously another portion of the solution with lemon-juice, until the blue colour appears distinctly red; this also serves to colour paper. The red test-paper is used for the purpose of recognising a class of substances opposed to acids—that is, alkaline or basic bodies; these restore the original blue colour of the paper, as can be seen by bringing it into contact with lime-water or solution of caustic potash.

Experiment 5.—Add gradually, with constant agitation, to one ounce of cold water, powdered saltpetre (potassium nitrate), as long as it continues to be dissolved—perhaps about a quarter of an ounce; if more is added than is necessary, it will remain undissolved at the bottom of the vessel. This solution is called a *cold saturated* solution. If this mixture be boiled, and saltpetre again be added, then about three ounces more will be required to saturate the water. A thermometer held in this *boiling saturated* solution will indicate about 240° F. (115·5° C.), while simple boiling water indicates only 212° F. (100° C.). *All saline solutions boil and freeze with more difficulty than water.* All bodies soluble in water behave in a similar manner—that is, they are soluble in it only in definite quantities, and in most cases hot water dissolves more of them than cold.

Experiment 6.—If the solution obtained in the last experiment be poured into a porcelain dish, previously heated, and be suffered to remain quiet until cold, then the three ounces of saltpetre which were last added separate again, not as a powder, but as regularly formed prisms. These prisms are six-sided, and are surmounted by two faces similar to a roof; they are called *crystals of saltpetre.* (Fig. 27.) All crystals are characterised by having planes, edges and angles, constructed, as it were, of simple triangular, quadrangular, or poly-angular pieces, with polished surfaces. This symmetry is found even in the interior of them, as can easily be seen by holding a transparent crystal towards the light, and turning it slowly round; or breaking it, when the fragments will often exhibit the same regular form which characterised the whole crystal. Thus, in inanimate nature, a mysterious power exists, similar to that which compels the bees to construct six-cornered cells, and the potato to produce its five-angled corolla and five stamens, and by which the smallest particles of bodies, called *molecules*, are forced to arrange themselves in a fixed order, assuming a regular shape. But this can in general only be accomplished by a body in its fluid or aeriform state, since *free motion* of the molecules is essential. Time also is required for this operation; hence crystals are always more regular the more slowly they are formed. Many of the splendid crystals which are dug from the depths of the earth were, perhaps, thousands of years in forming.

Fig. 27.



Experiment 7.—Evaporate the liquid remaining above the crystals in the former experiment at a gentle heat, until a scum is formed on the surface, then remove it from the fire and let the solution cool, *stirring it constantly* with a wooden stick. In this way, instead of crystals, a *powder of saltpetre* will be obtained consisting of minute crystals.

The liquid just alluded to may be regarded as a cold saturated solution, containing about a quarter of an ounce of saltpetre. If, by evaporation only, so much water is left as is sufficient when hot to keep in solution but a quarter of an ounce of the saltpetre, then crystals begin to appear in the form of a film on the cooler surface, indicating the saturation of the liquid. If this again is allowed to cool quietly, a second crop of crystals will be obtained; but, by continual stirring, they are broken at the moment of their formation—by slow movement into a coarse, and by rapid movement into a fine powder. This may be called *interrupted crystallization.* Sugar presents a similar example; the same syrup, which, if cooled quietly, yields sugar-candy, when stirred, yields common loaf-sugar.

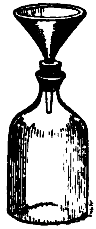
Experiment 8.—Put into boiling water as much common salt as will dissolve, and let the solution cool; no crystals will form, because salt is as soluble in cold as in hot water. Now evaporate one-half of the solution over a spirit-lamp, and set aside the other half in a warm place; in the first case, mere irregular grains of salt will be obtained, but in the latter case, after some days, regular cubes of salt will be deposited.

Experiment 9.—Dissolve a spoonful of salt and one of saltpetre in lukewarm water, and put the solution in a warm place, that the water may gradually evaporate; the two salts, which are intimately mixed in the solution, will upon crystallization separate completely from each other. The saltpetre separates into long prisms, containing no trace of the common salt, and the latter separates into cubes, entirely free from saltpetre. Thus the particles of salt and saltpetre did not attract each other; but, upon crystallizing out of the solution, the *homogeneous* salts assumed separately a regular form, precisely as if one only of these two substances had been dissolved.

PRESSURE.

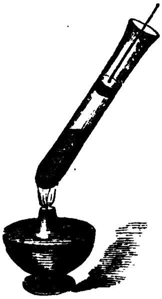
The earth is surrounded by air, as by a mantle; this mantle is called the *atmosphere,* and is supposed to extend about forty-five miles above the solid earth. The air possesses no colour, and is transparent; hence it is invisible, and its particles are so easily movable upon each other that it cannot be grasped by the hand. But it is rendered obvious that the air is material, and fills every space commonly called empty, by wrapping strips of moistened paper round a funnel, so that it may fit exactly into the mouth of a bottle; if the funnel be now filled with water, the fluid will *not* run into the bottle, as the air contained in the latter will not let it enter; but if the funnel be raised a little, the air escapes, and the water immediately rushes into the bottle. We learn also by the balance that a vessel apparently empty, *i.e.*, containing atmospheric air, weighs more than one which is really empty, as when the air has been exhausted from it. But air is so light that 800 measures of it weigh rather less than one measure of water, yet the atmosphere presses with great weight on the earth, and upon everything thereon. But this pressure is only noticed when the air is removed from a place, thus leaving it without counter-pressure. The total weight of the atmosphere has been estimated to be equal to that of a globe of solid lead sixty miles in diameter.

Fig. 28.



When a solid is immersed in a fluid (either a liquid or gas), it is pressed upon equally on all sides by the fluid with a force which depends partly upon the weight of the fluid and partly upon the depth to which the solid is immersed. A stone suspended one foot from the surface of water is only pressed upon with half the force that it would experience if sunk two feet; a stone sunk one foot in mercury would receive 13½ times as much pressure as it would if sunk one foot in water, because mercury is 13½ times as heavy as water. The same thing is true in regard to air and other gases, and bodies near the surface of the earth are pressed upon equally on all sides by the whole forty-five miles of air above.

Fig. 29.



Pressure of the Atmosphere.—Experiment 1.—Wrap some tow round one end of a stick and grease it with tallow, thus forming a plug, which must fit tightly into a strong test-tube. Boil some water in the test-tube, and when the air has been expelled by the steam, insert the plug; as the water cools the steam condenses, a vacuum is produced, and the plug will be pressed down upon the surface of the water; by heating, it is again forced up by the steam thus generated, and by immersing in cold water it is again forced down. In consequence of the cooling and condensation of the steam a vacuum is formed, and, therefore, the counter-pressure against the weight of the exterior air is removed; the pressure of the latter, accordingly, forces down the plug. On this principle the piston is forced up and down in the cylinder of many steam-engines.

This one-sided pressure often causes the ascent and reflux of liquids in tubes.

Experiment 2.—If water is boiled as was directed at page 39, by means of steam, and during the boiling the lamp is removed, then the pressure of the air acting on the surface of the water in the beaker-glass will very soon force the water contained in it through the tube back into the flask, which in a short time becomes quite filled with water. The counter-pressure of the steam must naturally decrease as it cools and condenses. As long as the lamp is under the flask, the pressure of the steam is stronger than that of the air, and the steam, being continually generated, forces the air previously contained in the flask into the water of the beaker-glass. This reflux of liquids is particularly to be feared, when such kinds of gases are conducted into water as are absorbed by it readily, and in large quantities. This is prevented by passing through the cork a second glass tube open at both ends, and letting it reach nearly to the bottom of the flask, by which tube air can penetrate into the flask as the pressure of the steam diminishes. This contrivance is called a ***safety-tube***.

Fig. 30.

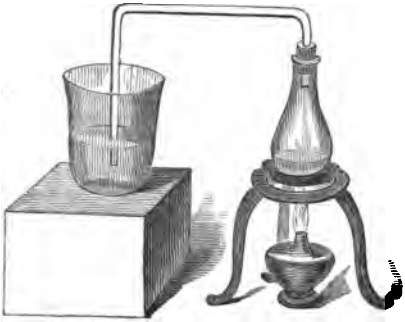
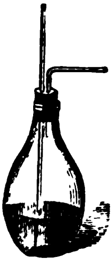


Fig. 31.



Barometer.—It can be proved by exact experiment that the atmosphere presses upon the earth with a force equal to that of a layer of mercury 30 inches deep, or a layer of water 13½ times deeper (34 feet), water being 13½ times lighter than mercury. The instrument by which the amount of atmospheric pressure can be observed and measured is called the *barometer*. Fill a glass tube, 32 inches in length, one end of which is closed, with mercury; close it with the finger, and invert it into a vessel of mercury; on removing the finger, the mercury will not run out, but will fall some inches, perhaps to *s* (Fig. 33). The height of the column of mercury, from *a b* to *s*, amounts to about 30 inches. The mercury does not fall lower, on account of the external pressure of the atmosphere, which is exerted on the mercury at *a b*, and not at *s*, since this end is closed. The column of mercury in the tube may be regarded as the counterpoise to the atmospheric pressure, and it is hence concluded that the latter exerts just as much pressure upon the earth as a column of mercury 30 inches high. If the tube be opened at the top, the pressure of the air on both extremities being then made equal, the mercury will flow from the tube. The space above the mercury, at *s*, is a vacuum, and is called the Torricellian vacuum, because it was first observed by Torricelli. In common barometers the tube is curved at the bottom, and provided with a bulb. This bulb is open at the top, and supplies the place of the vessel filled with mercury in the preceding figure. Here also the pressure is only exerted at one end, for the atmosphere can only press on the mercury contained in the bulb. The height from *o* (Fig. 34) to the top of the mercury amounts to about 30 inches.

Fig. 32.

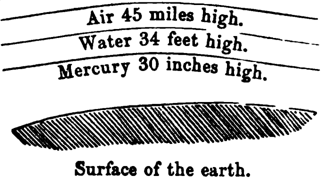
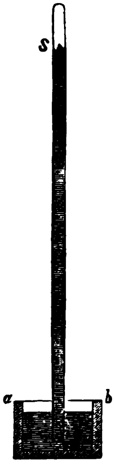
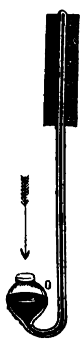


Fig. 33.



If weights be placed in one pan of a balance, the opposite one will rise, but on their removal it will sink. The same thing happens with the barometer. Any increase in the weight or density of the air presses the mercury up, and the barometer rises; but any diminution of weight will make it fall. The height of the mercury may be read off by fixing to the upper part of the tube a scale divided into inches and tenths of an inch. The mean state of the barometer is at 30 inches, and 31 is called a very high, and 29 a very low, state of the barometer. In this country, as a general rule, the north and east winds cause the barometer to rise, and the south and west winds cause it to fall. The former winds, blowing chiefly from the land, are cooler, and at the same time drier, than the latter, which pass over the ocean, there becoming saturated with moisture; the former likewise come from colder into warmer, while the latter, on the contrary, proceed from warmer into colder regions; by which the capacity of saturation for vapour is increased in one case and diminished in the other. Hence it is natural that, when north and east winds prevail, it should rain less frequently than during south and west winds; and that the former winds are dry, while the latter are damp.

Fig. 34.



Why water does not flow from a jar inverted over the pneumatic trough, why it continues to flow through a syphon after the air has been exhausted, why liquids will not run into a vessel when the air is confined, or why water will only rise to the height of 34 feet in a suction pump, are questions that scarcely require further explanation.

Elasticity of Gases.—All gases are perfectly elastic; they all contract equally under an increase of pressure, and expand when it is diminished. When a gas is collected in a bottle or other vessel the weight of it that the bottle will hold depends upon the pressure of the air at the time, which, as we have seen, is measured by the barometer. It is therefore of the utmost importance in measuring a gas to take account of the state of the barometer at the time. One hundred grains of air, for example, will occupy 333 cubic inches when the barometer is at 29, 322 when it is at 30, and only 311 when it is at 31 inches. As the alterations in volume are exactly in inverse proportion to the alterations of pressure, it is easy, after noting the volume of the gas and observing the height of the mercury in the barometer, to calculate what the volume would be at some standard pressure. The standard pressure at which gases are always measured is represented by 30 inches (or 760 millimetres) of mercury. The calculation required for this purpose is called the "correction for pressure." It must be applied, in addition to the correction for temperature (page 34), whenever a gas is to be measured. It can always be made by the following formula:

$$\frac{\text{Vol. at observed pressure} \times \text{Observed pressure}}{\text{Standard pressure}} = \left\{\begin{array}{l}\text{Volume of gas at}\\ \text{standard pressure.}\end{array}\right.$$

For example, we have 333 cubic inches of air when the barometer stands at 29 inches. How many cubic inches would there be if the barometer stood at 30 inches?

Here we must multiply by the observed pressure 29, and divide by the standard pressure 30:

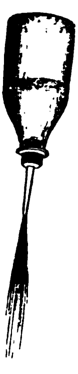
$$\frac{333 \times 29}{30} = 322 \text{ cubic inches (very nearly).}$$

The ordinary *air pump* depends for its action in producing a vacuum upon the elasticity and pressure of the air.

If the pressure or tension of a confined quantity of air be increased, by compressing it either directly or by the addition of more air, it can be forced to stream out from a small opening with great rapidity, as is shown on a small scale in the common bellows, and on a larger scale in the blacksmith's bellows. Should there be water before this opening, the air will press it out in a jet or stream.

Experiment 3.—Take a piece of a fine glass tube, drawn out to a point, and adapt it, by means of a perforated cork, to a bottle. Fill the bottle half full of water, and blow into it through the point of the tube; when the blowing ceases, the air will escape in a stream. But if the bottle be inverted as soon as the air is blown in, then the water will be spurted out by the compressed air above. Such an apparatus (the washing-bottle) is frequently employed for washing

Fig. 35.

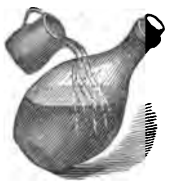


residues or precipitates remaining on filters, in order to free them from soluble matter. There is a similar contrivance connected with the common fire-engine, called the air-vessel, enabling it to throw an uninterrupted stream of water.

The pressure of the atmosphere exerts a great influence *on the boiling* of water, and of other liquids. If water is made to boil when the mercury in the barometer is very low (in foul weather), brisk ebullition will take place at about 210° F. (99° C.); when the mercury is very high (in clear weather), boiling will not occur under 214° F. (101° C.).

Experiment 4.—Heat a flask half filled with water till the water boils briskly; then remove it from the fire and quickly cork it; the boiling immediately ceases, but will commence again if cold water be poured over the upper part of the flask. In this manner it can be made to bubble or boil, even though it be only lukewarm. There is no air in the flask, it having been expelled by the steam, and it could not re-enter it, on the cooling and condensation of the steam, on account of its having been closed. Consequently there is no pressure of air on the water, and it will boil even at a temperature of 68° F. (20° C.). The boiling ceased on account of the pressure of the steam upon the water; but the steam being condensed by the cold water, the pressure was so much diminished, that a portion of water again became aeriform with a boiling motion. In many manufactories, an appropriate apparatus (vacuum-pan) has been contrived for boiling and evaporating in a vacuum, as, for instance, in sugar-houses.

Fig. 36.



The air is densest at the level of the sea, and thinner in proportion to its distance from the earth, as there is less air above it. Hence the mercury will stand lower, and water boil more easily, on the top of a mountain than in the valley below. On the top of Mont Blanc mercury rises only to the height of 16 inches in the barometer, and water boils at 183° F. (84° C.). Hence the barometer and the boiling point of water may be employed for calculating the heights of mountains.

As water boils more easily under diminished pressure, so *it boils with more difficulty when the pressure is increased.* An increase of pressure can be produced, not only by the air, but by the steam of the water itself, if new steam be constantly generated, while the escape of that already formed is prevented. This is best done by heating water confined in a strong and firmly closed vessel. For this purpose a Papin's Digester may be used, in which water may be heated to the temperature of 392° F. (200° C.), and, indeed, still higher, whilst in open vessels it cannot be heated above 212° F. (100° C.). If the amount of steam in it is twice as much as in an uncovered vessel, the pressure is said to amount to two atmospheres; if there is 3, 4, 5, 10, 20 times the quantity, there is said to be a pressure of 3, 4, 5, 10, 20 atmospheres. Vessels of this kind are often employed to effect a complete penetration of the water into solid and hard substances. Thus, for example, water at 212° F. (100° C.) dissolves the gelatinous matter only on the surface of the bones, whilst water at a temperature ranging from 230° F. (110° C.) to 248° F. (120° C.) entirely penetrates the bones, and extracts the gelatine also from the interior of them.

CHAPTER III.

LAWS OF CHEMICAL FORCE.

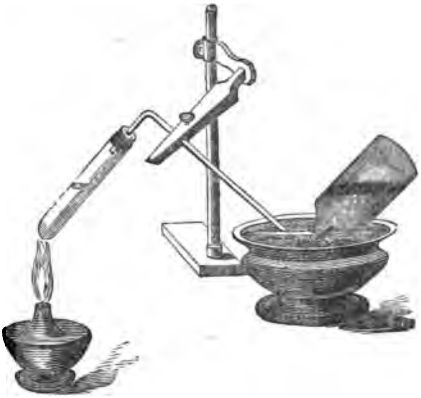
We have already, in Chapter I., seen something of the operations of that form of force which is known as chemical force. We have seen that by its operation *compounds* are produced by the union of *elements*, the compounds being in all cases different from the elements which compose them. And we have seen that the science of chemistry is mainly concerned with the changes which are brought about by chemical force. It is now necessary, before entering on the systematic study of chemical facts, to consider the *theories* of chemistry—the laws according to which chemical actions take place. These laws have been discovered by thousands of difficult and laborious investigations. For the most part it will not be possible for the student to prove them to himself experimentally: the processes necessary for this are too difficult, and require apparatus of too expensive a kind. He must be content in the outset of his study to take the laws upon trust.

MODES IN WHICH CHEMICAL ACTION TAKES PLACE.

Combination.—Experiment 1.—The simplest way in which chemical action can take place is when two substances, elements or compounds, combine directly together and produce a single new compound. Examples of this have already been given at page 13, but another may be added here. Take a globule of mercury about the size of a very small pea, heat it in a test-tube until it boils, and then throw in about an equal weight of iodine. Shake the tube slightly, so as to bring the two into contact. They will combine, and a beautiful scarlet compound called mercuric iodide will be found on the sides of the glass.

Decomposition.—Experiment 2.—Another kind of chemical change is called *decomposition*. It consists in a separation, either partial or complete, of the constituents of a compound from one another. The following experiment will illustrate it. Introduce into a somewhat tall, but not too thin, test-tube, 108 grains of mercuric oxide. One end of a bent glass tube is adapted to it by means of a perforated cork, and the other end is conducted into a vessel filled with water. Either suspend the tube by means of a piece of string or wire, or support it by a retort-holder. A retort-holder is a wooden stand, provided with a movable vice, by which glass vessels can be held in the most convenient manner, as shown in the annexed figure. Then heat the test-tube until all the mercuric oxide has disappeared. The red powder becomes black as the heat increases, and bubbles of air escape, which are collected in a glass bottle held over the end of the tube, this bottle having been previously filled with water and then inverted into the bowl, after closing the mouth of it with the finger or a glass plate. No water will escape until bubbles of air from the tube are passed into it, which, on account of their greater lightness, ascend and displace the water. When the water is displaced, remove the bottle and close it with a cork, replacing it by another bottle, likewise previously filled with water, and repeat this process until the evolution of gas ceases. The first bubbles that pass over consist of the atmospheric air contained in the test-tube, and expanded by the heat, but *oxygen gas* quickly succeeds. This is *one* of the component parts of mercuric oxide, and can easily be recognised by the vivid combustion in it of a glowing

Fig. 37.



shaving. At the same time there is formed on the upper part of the test-tube a brilliant metallic mirror, which consists of mercury, the *second* element of the red oxide. When the latter has entirely disappeared, *immediately* withdraw the tube from the water, let the test-tube cool, and unite the mercury adhering to its sides into a single globule, by means of a feather. It will amount in weight to 100 grains; this, subtracted from the original weight, 108 grains, leaves 8 grains, the amount of the oxygen. The red powder consists of a brilliant heavy metal and of a gas, two entirely dissimilar bodies, and it has been *decomposed* by the heat. If these are chemically combined together by proper means, they will unite again and form the red oxide, a body in which the peculiar properties of the mercury as well as of the oxygen are entirely lost.

Double Decomposition.—Experiment 3.—It very often happens that two compounds are capable of decomposing one another, thus producing two new compounds. This kind of chemical action is called *double decomposition.* Weigh out 27 grains of mercuric chloride (corrosive sublimate) and 33 of potassium iodide. Dissolve them separately in water and then mix the solutions. A yellow precipitate will appear, which in a few minutes will change to a beautiful scarlet powder, which will soon subside and leave the liquid colourless. The red powder is *mercuric iodide*, the same compound which was formed in Experiment 1. The clear liquid on evaporation will yield cubical crystals of *potassium chloride.* Mercuric chloride consists of the elements mercury and chlorine; potassium iodide, of potassium and iodine. When mixed, the mercury of the one unites with the iodine of the other, and the chlorine of the one with the potassium of the other.

Substitution, or Displacement.—Experiment 4.—Take a cold saturated solution (see page 50) of mercuric chloride, and immerse in it a strip of bright copper. Very soon the surface of the copper will become covered with a film of mercury, and in the course of a few hours the whole of the mercury will be separated in this way, and some of it may be collected into a globule at the bottom of the liquid. But in the mean time a portion of the copper will have been dissolved, and has now taken the place formerly occupied by the mercury, and the solution is found to be green. The copper has, in fact, *displaced*, or been *substituted for* the mercury in the original compound, so that we have now in solution copper chloride instead of mercuric chloride. For every 100 grains of mercury which are separated in the metallic form, 31 grains of copper are dissolved, so that we learn that 31 parts of copper will displace 100 of mercury.

When the whole of the mercury has been separated, the clear green solution may be poured into another glass and the bright blade of a knife immersed in it. The iron will displace the copper, just as the copper displaced the mercury, an iron chloride (called ferrous chloride) remains in solution, and the copper is precipitated on the unaltered iron, giving it at first the appearance of being converted into copper. The 31 parts of copper which displaced 100 of mercury are themselves displaced by 28 of iron. Examples of this kind are very common in chemistry.

QUANTITIES CONCERNED IN CHEMICAL CHANGES.

It will readily be understood from the foregoing illustrations that it is of the utmost importance in studying chemistry to attend to the *relative quantities* of different substances which enter into chemical changes. It is to the careful study of these quantities that the dignity of chemistry as an exact science is due.

Law of constant composition.—This, the fundamental law of chemistry, may be stated in these words:

A given chemical compound always contains the same elements in the same proportions.

The law scarcely requires explanation. Water always contains, in 9 parts by weight, 8 of oxygen and 1 of hydrogen; mercuric oxide always contains 8 of oxygen to 100 of mercury; mercuric iodide, 100 of mercury to 127 of iodine, and so on. If 100 grains of mercury are heated with 130 grains of iodine, 227 grains only of mercuric iodide are produced, and 3 grains of iodine remain unchanged.

Difference between combination by weight, and by volume.—Combination by volume means combination by measure, and it will be obvious to every one that, as bodies differ very much in weight, there will generally be a great difference between the proportions by measure and the proportions by weight in which two bodies combine. Take the case of the two gases hydrogen and chlorine (both elements). One volume or measure of hydrogen will combine with one volume of chlorine; but, as chlorine is 35½ times heavier than hydrogen, one part by weight of hydrogen really combines with 35½ parts by weight of chlorine.

In the case of solids, and to a great extent of liquids also, we are compelled to trust almost entirely to the evidence of weight, for solids are very difficult to measure, and both solids and liquids exhibit such irregularities in regard to the volume that they occupy, that it has not been found possible to frame constant laws for them. But with gases it is not so. They are very much easier to measure than to weigh; and as they exhibit a remarkable simplicity in the proportions by volume in which they act on one another; as moreover they are all equally affected in volume by alterations of temperature and pressure (Chap. II.), they are always measured for the purposes of experiment. Their weight (or specific gravity) is determined once for all with the utmost care, and it is then easy to calculate in a few minutes the weight of any given volume of any one of them.

CHEMICAL ACTION BETWEEN GASES.—CONSTITUTION OF GASES.

*When chemical action takes place between two gases, the volumes of those gases which are concerned in the change are found to bear a definite and generally a very simple proportion to one another, and also to the volume of the resulting gas, if the resulting compound be a gas.**

This fundamental law, which is sometimes called the "law of gaseous volumes," will be better understood with the aid of a few illustrations. In studying them it is necessary to bear in mind the fact mentioned on page 37, that all true vapours are gases. In considering the laws of gases, we have to include water-gas (steam), sulphur-gas, and even zinc-gas, for zinc is converted into gas at a high temperature, as well as oxygen and hydrogen, which are always gases.

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* It will be understood hereafter that the apparent exceptions to this rule are in perfect accordance with it. They only occur when a compound gas of complex structure is decomposed during a chemical change.

We deal in fact with matter in the gaseous state. For the sake of exact comparison all gases and vapours are assumed to be measured at a temperature of 32° F. (0° C.) and a pressure of 30 inches (760 millimetres) of mercury. Even if, as is the case with steam, the gas cannot really be cooled to 32° without condensing, it is easy to calculate by the formula already given (page 34) what its volume would be if it could.

1 volume of chlorine gas combines with 1 volume of hydrogen gas to form 2 volumes of hydrochloric acid gas. This is the simplest case.

2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of water-gas, or steam. Here 3 volumes condense, during the act of combination, into 2.

1 volume of arsenic vapour combines with 3 volumes of oxygen to form 1 volume of the vapour of white arsenic (arsenious anhydride). Here 4 volumes become condensed into 1.

These three cases will serve to illustrate the great simplicity that exists in the proportions in which gases unite by measure. Any given measure, pints, cubic inches, or litres, may of course be substituted for volumes in the above examples, and we may, for instance, say that every 2 pints of steam contain 2 pints of hydrogen and 1 pint of oxygen condensed into 2.

But this is not all. When a gas is decomposed into other gases, the volumes of its separated constituents are all found to bear simple and definite proportions to that of the original gas.

The gas called nitrous oxide, or laughing gas, is composed of nitrogen and oxygen, but it cannot be formed by the direct combination of these two elements. Nevertheless it is found, that when it is decomposed every 2 volumes yield 2 volumes of nitrogen and 1 volume of oxygen, so that there can be no doubt that its constitution is similar to that of steam.

Ammonia gas, again, is composed of nitrogen and hydrogen. When it is decomposed by electric sparks every 2 volumes yield 1 volume of nitrogen and 3 of hydrogen. The gases when separated from one another occupy in fact twice the volume that they did in a state of combination.

Standard of Volume for Gases.—The next point that we have to determine is, how best to represent the constitution of the numerous gases which are known to us; or in other words, how to represent the kind and the quantity of the elements which are present in equal volumes of them. It is obviously necessary, if we would compare them, to deal with equal volumes. It would appear at first sight that the best and simplest plan would be to state in all cases the composition of *one volume* of each gas. It is quite possible to do so, and indeed it does not very much matter what volume we take as the standard. But if we select one volume we are compelled to make constant use of fractions in representing the composition of gases, and it is therefore usual to take *two volumes* as the standard of comparison. A single example will make this plain.

We have already seen that 1 volume of chlorine combines with 1 volume of hydrogen to form 2 volumes of hydrochloric acid. If we wish to state the composition of 1 volume of hydrochloric acid, we must do so by saying that it contains ½ a volume of chlorine and ½ a volume of hydrogen. But on the 2 volume system the fractions are avoided. 2 volumes of chlorine are said to combine with 2 of hydrogen to form 4 of hydrochloric acid; and 2 volumes of hydrochloric acid contain 1 of chlorine and 1 of hydrogen.

Let us therefore agree to state in every case the composition of 2 volumes of gas. If we do so, we must never represent less than 2 volumes as taking part in any chemical process. We must in fact assign arbitrarily to the smallest quantity of any gas that is ever concerned in any chemical change a volume of 2. It is almost needless to add that our facts will then be equally true if we apply them to 2 millionths of a cubic inch or to 2 gallons.

Chemical Symbols and Formulæ.—Each one of the elements is denoted in the system of modern chemistry by a *symbol*, which is either the first letter, or the first and some other characteristic letter, of its *Latin* name. Thus, oxygen is denoted by O; hydrogen by H; carbon by C; chlorine, which begins with the same letter, by Cl; calcium by Ca, and so on. In most cases the Latin name is the same as the English, but in a few instances it is different and the symbol must be remembered carefully. The following are the exceptions: Antimony is Sb., from stibium; copper, Cu., from cuprum; gold, Au., from aurum; iron, Fe., from ferrum; lead Pb., from plumbum; mercury, Hg., from hydrargyrum; potassium, K., from kalium; silver Ag., from argentum; sodium Na., from natrium; tin, Sn., from stannum; and tungsten, W., from wolfram. Compounds are described by *formulæ*, which consist of the symbols of the elements composing them. Thus the formula for hydrochloric acid is HCl., for mercuric oxide, HgO., and so on.

Symbols then are employed to denote the elements, but they are also employed, by a very useful extension, to denote *definite weights of them*. What these weights are we have now to consider, confining ourselves for the present exclusively to elements in the state of gas.

To begin with, let us represent the standard 2 volumes of hydrogen by the symbol H, and 2 volumes of chlorine by the symbol Cl. How are we in this case to represent hydrochloric acid, which, as we have already seen, contains in 2 volumes, 1 of hydrogen and 1 of chlorine? 2 volumes of this compound must evidently be represented by such a formula as this: $H\frac{1}{2}$ $Cl\frac{1}{2}$. This is inconvenient, and yet such cases will be of incessant occurrence, unless we alter our representation of the elements. Accordingly we describe the 2 volumes of hydrogen as H_2 and the 2 volumes of chlorine as Cl_2, and the formula for hydrochloric acid, which contains $\frac{1}{2}$ of each, then becomes HCl.

To take another instance. The symbol N might be applied to 2 volumes of nitrogen, if we did not know that 2 volumes of ammonia contain only 1 volume of nitrogen united to 3 volumes of hydrogen. If we use the symbol N for 2 volumes of nitrogen and H for 2 volumes of hydrogen, ammonia must be described by the inconvenient formula $N\frac{1}{2}$ $H1\frac{1}{2}$. But calling nitrogen N_2, and hydrogen H_2, both of these are doubled, and the formula for ammonia becomes NH_3.

Again, with phosphorus: we might apply the symbol P to 2 volumes of phosphorus vapour if it were not for such compounds as phosphine gas, 2 volumes of which contain only one quarter as much phosphorus as is present in 2 volumes of phosphorus vapour. But by describing phosphorus as P_4 and phosphine as PH_3, we get over the difficulty; for the formulæ show plainly, and in whole numbers, that one contains four times as much phosphorus in 2 volumes as the other.

We thus gain more definite ideas of the signification which may be applied to the symbols and formulæ of gases. For the sake of distinctness they may as well be put in the form of definitions. :—

1. *The symbol of an elementary gas is a letter or letters, used to denote the smallest fraction of the weight of the normal two volumes of it that is ever found in two volumes of any compound gas.*

2. *The formula, whether of an elementary or compound gas, must always exhibit the composition of two volumes of it. It consists of symbols.*

Thus, the symbol for hydrogen is H; the formula, H_2; the symbol for phosphorus is P; the formula, P_4. Compounds must always be described by formulæ, the term "symbol" being reserved for elements.

Relative Weight of Gases.—It has already been shown (page 23) that the specific gravity of a body means its weight as compared with the weight of an equal volume of some other body which is taken as a standard. The standard for gases is hydrogen, one volume of which is said to weigh 1. We have therefore only to double the specific gravities of the gases (which have been carefully determined by experiment) to find the relative weights of the standard 2 volumes of each gas. If 2 volumes of hydrogen weigh 2 (2 grains, 2 pounds, or 2 hundredweight), 2 volumes of chlorine will weigh 71, 2 volumes of oxygen, 32, and so on with the rest. We must now study the relative weights of the different elements which go to make up the weight of 2 volumes of each of the more important gases, elementary and compound. It will be best to take the elementary gases first.

Constitution of Elementary Gases.—It has already been seen that the formula for 2 volumes of any elementary gas is determined by the compounds which contain that gas. The formula for hydrogen is H_2, because compounds are known which contain in 2 volumes only half as much hydrogen as is contained in 2 volumes of the pure element. Now what is true of a large volume must also be true of any volume, however minute, and we are therefore led to the conclusion that even the smallest conceivable volume of hydrogen gas that exists in a separate state must be capable ot division into two parts, and as the weight of 2 volumes of hydrogen has already been defined to be 2 (hydrogen being the standard), each of these constituent parts must have a weight of 1.

*Atoms.**—The name *atom* is, for reasons that will be explained hereafter, applied to the quantity of each elementary gas that is denoted by its symbol. 2 volumes of hydrogen denoted by the formula H_2, weigh 2, and consist of 2 *atoms*, each denoted by the *symbol* H, and each weighing 1. In like manner, 2 volumes of phosphorus vapour are said to consist of 4 atoms; the formula for 2 volumes being P_4 and the symbol for each atom P. The atom in this case weighs 31. In the case of mercury and a few other elements the 2 volumes of gas are said to contain but 1 atom (weighing, in the case of mercury, 200), because no gas is known which contains in 2 volumes less than 200 of mercury.

We can now add the definition of an atom to those given above:

3. *The atom of an elementary gas is the quantity denoted by its symbol; that is, the smallest fraction of the weight of 2 volumes of it that is ever found in 2 volumes of any other gas.*

These remarks will sufficiently explain the following table (p. 70) which exhibits the composition of all the more important elementary gases. It is only necessary to point out that two of them, oxygen and sulphur, occur in two places. Two different modifications of each of these gases are known, which differ from one another in weight. The *atoms* are, however, the same in each case.

Constitution of Compound Gases.—It will now be obvious that two volumes of every compound gas must contain two or more atoms of two or more kinds. The formula for a compound gas must denote the number and kind of atoms that there are in two volumes of it. Hydrochloric acid is an example of the simplest kind of gaseous compound, two

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* I use this word in its present connection with regret. It would certainly be better to exclude it altogether from an account of gases which is independent of the truth or untruth of the atomic theory. But there is no word in use which exactly answers to the conditions, and I hold the coinage of a word to be far too serious an experiment to be undertaken in an elementary treatise, or by any but a leader in science. Might not the word "prime," originally employed by Wollaston, be conveniently revived for this purpose?

ELEMENTARY GASES.

Name of Gas.	Sp. Gr. or weight of 1 volume.	Weight of 2 volumes.	No. of atoms in 2 volumes.	Symbol for each atom.	Weight of each atom.	Formula for 2 volumes.
Mercury	100	200	1	Hg.	200	Hg.
Zinc	32·5	65	1	Zn.	65	Zn.
Cadmium	56	112	1	Cd.	112	Cd.
Hydrogen	1	2	2	H.	1	H_2
Chlorine	35·5	71	2	Cl.	35·5	Cl_2
Bromine	80	160	2	Br.	80	Br_2
Iodine	127	254	2	I.	127	I_2
Oxygen	16	32	2	O.	16	O_2
Sulphur (above 1000° C.)	32	64	2	S.	32	S_2
Nitrogen	14	28	2	N.	14	N_2
Oxygen (as ozone) . .	24	48	3	O.	16	O_3
Phosphorus	62	124	4	P.	31	P_4
Arsenic	150	300	4	As.	75	As_4
Sulphur (below 1000° C.)	96	192	6	S.	32	S_6

volumes of it containing, as we have already seen, one atom of hydrogen and one of chlorine. Other gases contain atoms of three and even four different elements, and there are sometimes as many as twenty, thirty, or even more atoms in two volumes. In fact, we know no limit to the number of atoms which may enter into the composition of a compound gas.

The following table (p. 71) shows the constitution and formulæ of a few of the more important compound gases. The rules according to which they are named will be explained hereafter.

Compound Gases containing non-volatile Elements.—Many gases contain atoms of elements which have not as yet been converted into the condition of gas, or which, even if they do exist as gas, have never had their specific gravities accurately determined in that condition. In these cases the atomic weight (weight of the atom) of the non-volatile element can be found by exactly the same rule as that given before, namely, by observing the smallest weight of the element that ever enters into the composition of 2 volumes of a compound

COMPOUND GASES.

NAME OF GAS.	Specific gravity.	Weight of 2 volumes.	Kind, number, and weight of the atoms in 2 volumes.	Formula.
Mercuric Chloride .	135·5	271	Mercury 1 = 200 Chlorine 2 = 71	$Hg\,Cl_2$
Hydrochloric Acid .	18·25	36·5	Hydrogen 1 = 1 Chlorine 1 = 35·5	$H\,Cl$
Hypochlorous Anhydride. . .	43·5	87	Oxygen 1 = 16 Chlorine 2 = 71	Cl_2O
Water. (Steam) .	9	18	Hydrogen 2 = 2 Oxygen 1 = 16	H_2O
Hydrogen Sulphide.	17	34	Hydrogen 2 = 2 Sulphur 1 = 32	H_2S
Sulphurous Anhydride . . .	32	64	Oxygen 2 = 32 Sulphur 1 = 32	$S\,O_2$
Ammonia . . .	8·5	17	Hydrogen 3 = 3 Nitrogen 1 = 14	H_3N
Nitrous Oxide . .	22	44	Nitrogen 2 = 28 Oxygen 1 = 16	N_2O
Nitric Oxide . .	15	30	Nitrogen 1 = 14 Oxygen 1 = 16	$N\,O$
Phosphine . . .	17	34	Hydrogen 3 = 3 Phosphorus 1 = 31	H_3P
Phosphorous Chloride	68·75	137·5	Chlorine 3 = 106·5 Phosphorus 1 = 31	Cl_3P
Arsine	39	78	Hydrogen 3 = 3 Arsenic 1 = 75	H_3As

gas. But the definition of such an atom must be slightly modified, because as we do not know the specific gravity of the elementary gas, we cannot know the weight of 2 volumes of it, and consequently cannot know what fraction of that weight the atom is. For such atoms the following definition is accurate:

4. *The atom of a non-volatile element* (or indeed of any element) *is the smallest weight of that element that is ever found in two volumes of any gas.*

Take the case of carbon, an element which has never been converted into vapour. Many gases containing carbon are known, but not one of them contains in 2 volumes less than 12 of carbon. The atomic weight of carbon is therefore taken to be 12 and the atom is denoted by the symbol C., but not knowing the weight of 2 volumes of carbon vapour we cannot tell what fraction 12 is of that weight; or, in other words, how many atoms of carbon there are in 2 volumes of carbon gas.

It is obvious that we cannot have a *formula* for a non-volatile element, and that even in the case of volatile elements it is incorrect to apply the formula belonging to the gas to the liquid or solid element. For liquid and solid elements we must be content to use the *symbol* which denotes the atom.

The following table (p. 73) will show the constitution of a few important gases of the kind here specified and also the way in which the atomic weights of their non-volatile constituents can be determined from them.

The atomic weight of iron, as deduced from the composition of the vapour of ferric chloride, is 112. But ferric chloride is so similar to chromic chloride that the atomic weight of iron is held to be 56. The formula for ferric chloride is then $Fe_2\,Cl_6$; 2 atoms of iron, each 56, united with 6 atoms of chlorine, each 35·5. This brings it into analogy with chromic chloride $Cr_2\,Cl_6$. The composition and properties of other iron compounds confirm this view.

Calculation of the Specific Gravities of Gases from their Formulæ.—If the formula for any gas be accurate; that is, if it truly represents 2 volumes of the gas, the specific gravity of the gas may be calculated from the formula. This may be illustrated by an example.

The formula for carbonic anhydride gas is CO_2. *What is its specific gravity?*

The formula tells us that 2 volumes of the gas contain 1 atom of carbon, the weight of which we know to be 12, and 2 atoms of oxygen, each of which weighs 16, total 44. 2 volumes of carbonic anhydride therefore weigh 44, whereas 2 volumes of hydrogen weigh 2; which shows that carbonic anhydride is 22 times heavier than hydrogen; or, in other words, that its specific gravity is 22.

To find the specific gravity of any gas, elementary or compound, we have therefore only to find the weight of 2 volumes of it, by adding together the weights of the constituent atoms, and then to divide the number so obtained by 2.

COMPOUND GASES CONTAINING NON-VOLATILE ELEMENTS.

NAME OF GAS.	Specific gravity.	Weight of 2 volumes.	Kind, weight, and number of the atoms in 2 volumes.	Formula.	Symbol and atomic weight of the non-volatile constituent.
Methene .	8	16	Hydrogen 4= 4 Carbon 1= 12	CH_4	C = 12
Ethylene .	14	28	Hydrogen 4= 4 Carbon 2= 24	C_2H_4	
Ethene .	15	30	Hydrogen 6= 6 Carbon 2= 24	C_2H_6	
Carbonic oxide .	14	28	Oxygen 1= 16 Carbon 1= 12	CO	
Carbonic anhydride	22	44	Oxygen 2= 32 Carbon 1= 12	CO_2	
Cyanogen .	26	52	Nitrogen 2= 28 Carbon 2= 24	C_2N_2	
Chloroform	59·75	119·5	Hydrogen 1= 1 Chlorine 3=106·5 Carbon 1= 12	$CHCl_3$	
Silicon chloride	85	170	Chlorine 4=142 Silicon 1= 28	$SiCl_4$	Si = 28
Methyl silicate .	76	152	Carbon 4= 48 Hydrogen 12= 12 Silicon 1= 28 Oxygen 4= 64	$C_4H_{12}SiO_4$	
Chromyl chloride	77·75	155·5	Oxygen 2= 32 Chlorine 2= 71 Chromium 1= 52·5	CrO_2Cl_2	Cr = 52 5
Chromic chloride	159	318	Chlorine 6=213 Chromium 2=105	Cr_2Cl_6	
Ferric chloride	162·5	325	Chlorine 6=213 Iron 1=112	$FeCl_6$ . .	Fe = 112

COMPOSITION AS DETERMINED BY WEIGHT.—LAWS OF COMBINATION BY WEIGHT.

We must now leave on one side for the present all considerations of volume and confine ourselves to the examination of the *weights* of different substances which are concerned in chemical changes, and the constitution of compounds as determined by weight. This method of study is of universal application, and it is used indifferently for solids, liquids, and gases, whereas we have already seen that the study of volumes only gives satisfactory results when applied to gases.

Percentage Composition.—The most obvious way of stating the composition of a compound is as parts in 100. The results of an analysis are always calculated in this way first of all, the figures being usually carried to the second decimal place. Thus, if 20 grammes of lime are analysed they are found to contain 14·286 grammes of calcium, and 5·714 of oxygen. By a simple proportion sum it is then found that the percentage composition is:

Calcium	71·43
Oxygen	28·57
	100·00

For 20 : 14·286 :: 100 : 71·43
And 20 : 5·714 :: 100 : 28·57

The percentage composition of a few important hydrogen compounds is shown in the following table (p. 75).

Simplest numerical Proportion of the Constituents.—From the percentage composition, it is easy to calculate the proportion that the weight of each constituent bears to that of some one which is taken as unity. In the following table the last column shows the weights of various elements which are combined with 1 of hydrogen in a few important compounds, the proportion being of course the same as in the percentage composition. Any other element might be taken instead of hydrogen as the standard. It was indeed common at one time to take oxygen as the standard, calling it 100, but hydrogen is more convenient and is now generally employed.

When the composition of hydrogen compounds is represented in this way, very simple numbers are for the most part obtained, and some numerical relations are observed which are hidden in the mere percentage compositions. Thus in the three carbon compounds (which are examples of a large number actually known) we observe that 1 of hydrogen unites with 3, 6, and 12 of carbon. 3 is the smallest proportion ever found.

NAME OF COMPOUND.	Composition of 100 parts.		Composition. H = 1.
Hydrochloric acid .	Hydrogen	2·74	1
	Chlorine	97·26	35·5
Water	Hydrogen	11·1	1
	Oxygen	88·8	8
Ammonia . . .	Hydrogen	17·65	1
	Nitrogen	82·35	$4\cdot\dot{6}$
Methene	Hydrogen	25·	1
	Carbon	75·	3
Ethylene . . .	Hydrogen	14·29	1
	Carbon	85·71	6
Acetylene . . .	Hydrogen	7·69	1
	Carbon	92·31	12
Hydrogen sulphide.	Hydrogen	5·88	1
	Sulphur	94·12	16

By extending the above table we can easily obtain a series of numbers which represent the proportions in which a good many of the most important elements combine with one part of hydrogen. It is now necessary to extend our study to those compounds which do not contain hydrogen.

And here there are two courses open to us. We may begin by choosing some other element as a standard, calling its quantity 1, and calculating the weight of other elements which combine with it. But this method, though it serves to exhibit some interesting relations, does not bring out the simple general law towards which we are tending. We must make the comparison in another way.

Knowing the weight of any element which unites with 1 of hydrogen, let us calculate the weight of other elements which unite *with that weight.* For example: 35·5 of chlorine unite with 1 of hydrogen. Instead therefore of calculating how much of each other element will unite with 1 of chlorine, let us calculate how much will combine with 35·5 of chlorine. And inasmuch as 8 of oxygen combine with 1 of hydrogen, let us see what weight of each other element unites with 8 of oxygen. In examining the results so obtained, it must be remembered that two elements often unite in several different proportions.

Confining ourselves to the elements mentioned on the preceding table, we can obtain, from the analyses of well-known compounds, the following figures, which are merely samples of an immense number at our disposal:

35·5 parts by weight of chlorine unite with 1 of hydrogen; with 8, 24, and 32, of oxygen; with 3, 4, 6, and 12, of carbon, and with 32 of sulphur.

8 parts of oxygen unite with 1 and with 0·5 parts of hydrogen; with 8·875, $11{\cdot}8\dot{3}$, and 35·5 parts of chlorine; with 2·8, 3·5, $4{\cdot}\dot{6}$, 7 and 14 parts of nitrogen; with 3 and 6 parts of carbon, and with $5{\cdot}\dot{3}$ and 8 parts of sulphur.

$4{\cdot}\dot{6}$ parts of nitrogen unite with 1 of hydrogen; with $2{\cdot}\dot{6}$, $5{\cdot}\dot{3}$, 8, $10{\cdot}\dot{6}$, and $13{\cdot}\dot{3}$ parts of oxygen, and with 4 parts of carbon.

3 parts of carbon (the smallest quantity that unites with 1 of hydrogen) unite with 0·25, 0·5 and 1 part of hydrogen; with 8·875, 17·75, 26·625, and 35·5 of chlorine; with 4 and 8 parts of oxygen; with 7 of nitrogen, and 16 of sulphur.

16 parts of sulphur unite with 16 and 24 of oxygen; with 17·75 of chlorine and with 3 of carbon.

A careful study of these figures will elucidate two of the most important laws of chemistry. It will be seen:—

1. *That in some cases the proportions in which two elements combine with one of hydrogen are exactly the proportions in which they unite with one another.* 35·5 parts of chlorine and 8 parts of oxygen will respectively combine with 1 of hydrogen, and a compound is known which contains in every 43·5 parts, 35·5 of chlorine and 8 of oxygen.

2. *That even when two elements unite in proportions different from those in which they unite with one of hydrogen, the numbers representing the two proportions bear a simple relation to one another.*—Chlorine forms three compounds with oxygen. The one mentioned above contains 35·5 of chlorine to 8 of oxygen; the other two, 35·5 chlorine to 24 and 32 of oxygen.

In these two latter the proportion of oxygen is exactly three and four times as great as is found combined with hydrogen in water. The five oxides of nitrogen afford a still more striking example, although some thought is required to make it apparent. We already know that 1 of hydrogen unites with 4·6̇ of nitrogen. In the following table the columns next to the names of the compounds show the quantities of oxygen which are combined with 4·6̇ of nitrogen in each of the oxides. These quantities of oxygen are represented by somewhat awkward fractions. But if we multiply the 4·6̇ by 3, and note the quantities of oxygen which unite with 14 of nitrogen, as in the middle columns of the table, we find that the proportions of oxygen can be expressed by whole numbers, and furthermore that these numbers are simple multiples of 8, which we already know is the proportion in which oxygen unites with 1 of hydrogen. For the sake of comparison the percentage compositions are given in the last columns.

OXIDES OF NITROGEN.

					Parts in 100.	
	Nitrogen.	Oxygen.	Nitrogen.	Oxygen.	Nitrogen.	Oxygen.
Nitrous oxide	4·6̇	2·6̇	14	8	63·6	36·4
Nitric oxide	4·6̇	5·3̇	14	16	46·7	53·3
Nitrous anhydride	4·6̇	8·0	14	24	36·8	63·2
Nitric peroxide	4·6̇	10·6̇	14	32	30·4	69·6
Nitric anhydride	4·6̇	13·3̇	14	40	25·9	74·1

Units of Combining Weight.—Combining Weights.—Atomic Weights.—By an extension of the above methods of comparison, it is at last made out that a number may be found for each element which is variously known as its "combining weight," "proportional number," or, for reasons which will be explained hereafter, its "atomic weight."* This number denotes the smallest quantity which is ever found united with one of hydrogen, or with an analogous quantity of any other element. For shortness it may be spoken of as the *atom* of the substance. When we have once fixed upon the numbers which shall stand for the atoms of the elements, it is easy to represent the composition of all compounds by stating the number and kind of atoms which they contain. Thus with the oxides of nitrogen. If we agree that the atom of nitrogen is 14 and that of oxygen 8, we see by the above table that in the 5 oxides, one atom of nitrogen is united with 1, 2, 3, 4 and 5 atoms of oxygen.

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* Here, again, I think that the word "prime," used by Wollaston, might be conveniently applied. The expression 1, 2, or 3 primes is accurate, and involves no hypothesis; but to talk of 1, 2, or 3 combining weights or combining numbers, is not only inconvenient but absurd.

Use of Symbols.—The symbols which have already been used (page 68) to denote volumes of gas, may be applied with equal exactness of meaning to the atoms of elements as found by weight. Thus, H may denote 1 atom, or 1 part by weight of hydrogen. N, 1 atom, or 14 parts of nitrogen. O, 1 atom, or 8 parts of oxygen, and so on. With these symbols, formulæ, perfectly analogous to those already explained, may be constructed and applied to compounds. The formulæ for the 5 oxides of nitrogen are, if the above mentioned atomic weights be adopted: NO, NO_2, NO_3, NO_4, and NO_5.

Modes of fixing the Atomic Weights of Elements.—The numbers found for the combining or atomic weights of the elements from the study of the percentage composition of their compounds are liable to one serious drawback. We can never be certain that the number adopted represents the weight of any real unit of the element. We cannot indeed by fair reasoning conclude that there is any constant unit of combining weight for each element. All that we can say is, that if there be such a unit, it must either have the weight we have assigned to it or bear some simple numerical relation to that weight. Looking at the carbon compounds, for example, in the table on page 75, we come to the conclusion that the atomic weight of carbon is either 3, 6, or 12, for carbon unites with 1 of hydrogen in all those proportions. The atomic weight of nitrogen indicated by its hydrogen compound is 4·6, but the oxygen compounds of the same element seem to require it to be 14. In the same way the atomic weight of oxygen may be 4, 8, 16, or even some less simple number, and in all these cases analysis is powerless to tell us with any certainty which of the different numbers should be adopted. In fact, to speak accurately, analysis would appear to indicate that each element had more than one, or indeed many combining weights, but that these weights all bore a simple relation to one another.

Fortunately, however, there are other facts at our disposal with regard to the elements of their compounds which give clear evidence upon this point, and which not only indicate that there is a single unit of combination for each element, but enable us to fix its weight with a near approach to certainty. The chief of these exterior sources of evidence may be briefly mentioned in this place.

1. *The Law of Gaseous Volumes*, with the extensions of it that have already been considered (page 64, *et seq.*). This method can of course only be applied where gases are concerned, but its evidence when it can be obtained is more valuable than any other. The weights of the atoms of a great many of the elements are fixed in this way, and the method has been so fully explained that it is unnecessary to do more than refer to it here. We have seen, to take a single example, that the atomic weight of oxygen is taken as 16, because 16 is the smallest weight of oxygen that occurs in 2 volumes of any gas.

2. *Substitution.—Equivalents.*—This is merely a modification of the method by weight which has already been described. It has been seen (page 62) that one of the commonest modes in which chemical action takes place is by the *substitution* of a certain weight of one substance for another. In a series of substitutions we have only to ascertain the weight of different substances, which will displace one another to obtain a series of numbers which are called the *equivalents* of those substances. Starting from hydrochloric acid, for instance, which contains 1 of hydrogen to 35·5 of chlorine, we find we can replace the 1 of hydrogen by 23 of sodium, 39 of potassium, 20 of calcium, and so on. And in the same way the 35·5 of chlorine may be replaced (either directly or indirectly) by 8 of oxygen, 4·6 of nitrogen, 80 of bromine, 16 of sulphur, and so on. These various numbers are the *equivalents* of the elements, and it is found that the quantities they represent are not only equivalent to 1 of hydrogen, or 35·5 of chlorine, but also to one another. 16 of sulphur may take the place of 8 of oxygen or 80 of bromine, and 20 of calcium are always equivalent to 39 of potassium.

All this is simple enough, and it would seem as if the equivalents so found might be adopted as the most convenient of "combining weights." They were in fact adopted, and until lately were pretty generally employed for this purpose. But certain difficulties and inconveniences attend their use and have at last occasioned their abandonment. For, in the first place, the equivalents of some elements do not agree with the numbers indicated by the law of volumes. The equivalent of oxygen, for instance, is 8, but its atom, as deduced from the composition of compound gases, is 16. A still graver objection lies in the fact that as elements often combine in more than one proportion, it then follows that there are several equivalents for the same element. The case of the oxides of nitrogen, already cited, will make this plain. On page 76, it is stated that 8 parts of oxygen will combine with 1 of hydrogen, and with 2·8, 3·5, 4·6, 7 and 14, of nitrogen; it is therefore evident that these five quantities of nitrogen are all, in different compounds, equivalent to 1 of hydrogen, and that, in fact, nitrogen has five equivalents. This difficulty is indeed only another form of that which was pointed out at the beginning of this section.

But although the equivalent values of the elements are no longer thought sufficient to fix their atomic weights, this method of study is still of immense importance in chemistry, for two reasons:

Firstly, because, as we have seen, the true atomic weight, when not identical with the equivalent, always bears a simple relation to it.

And secondly, because it often enables to fix the exact constitution of a compound, as will be seen from the following simple illustration. The formula deduced from the analysis of acetic acid is CH_2O ($C = 12$, $H = 1$, $O = 16$). But we find by experiment that one quarter of the hydrogen of acetic acid may be replaced by a metal—by sodium, for instance, or silver. Now it is evident that, if the above formula were correct, the metal must displace *half an atom* of hydrogen, which, as the atom is defined to be the smallest possible quantity, is absurd. We must, therefore, double the formula and make it $C_2H_4O_2$. It is then seen to consist of 2 atoms of carbon, 4 of hydrogen, and 2 of oxygen. Sodium displaces one atom of hydrogen, and the formula for the compound so obtained is $C_2H_3NaO_2$. This view is completely confirmed by the law of volumes. By the rule given on page 72, the specific gravity of the vapour of acetic acid ought, if the formula CH_2O were correct, to be 15, whereas experiment shows it to be 30.

3. *Specific Heat.—Atomic Heat.*—This method gives very good results with elements in the solid state, and is of the utmost importance. Equal weights of different substances, heated to an equal extent, require very different times for cooling down to the same point. If one pound of water and one pound of mercury, at a temperature of 100° C. (boiling point) be allowed to cool down to 50°, it will be found that the water is 30 times as long in doing so as the mercury, and gives out 30 times as much heat. Water at any one temperature contains 30 times as much heat as an equal weight of mercury at the same temperature, and analogous differences are observed with all other substances. The quantities of heat which equal weights of different bodies contain at the same temperature (measured by the time required for their cooling, or in some other way) are called their *specific heats.* The specific heat of water is greater than that of any other known substance (except hydrogen), and it is therefore taken as the standard, and called 1. The specific heat of liquid mercury is said to be 0·033.

Now if we compare together the specific heat of different elements in the solid state, we find, as the following table (page 82) will show, that no relation can be observed between them. But if we compare together the specific heats, not of equal weights, but of *atomic weights,* which we can easily do by multiplying the specific heats of the first column by the atomic weights, we find that the numbers are very similar to one another, and approximate to 6·3. It is therefore inferred that if we could compare all elements in the solid state and at equal temperatures, in quantities proportionate to their atomic weights, and could avoid all errors of experiment, *they would all contain the same quantity of heat.* This quantity, found by multiplying the specific heat by the atomic weight, is called the *atomic heat* of the element.

It is easy to understand the assistance which this theory, discovered by Dulong and Petit (we must remember that it is only a theory, for its truth cannot be demonstrated), affords in fixing the atomic weights of elements. Let us take the case of zinc to illustrate it. The equivalent of zinc, as compared with hydrogen, is 32·5, and that number was formerly taken as the atomic weight of the metal. But the specific heat of zinc is 0·0955, and this number multiplied by 32·5 only gives 3·1, so that the atomic heat of zinc must be regarded as exceptional and only half of the usual quantity. But if we take 65 as the atomic weight of zinc, the anomaly vanishes, and the atomic heat becomes 6·2. The atomic weights of many other elements have of late years been doubled for the same reason.

Some few exceptions to this rule of atomic heat are known, of which carbon and silicon are the most important.

It will be seen that if we wish to find the atomic weight of an element from its specific heat, we must divide the average atomic heat 6·3 by the specific heat. The atomic weight so obtained will not be exact, but it will serve to indicate whether a particular number or a fraction or multiple of it is to be chosen as the true atomic weight.

E.g. Atomic weight of tin is $\frac{6\cdot3}{\cdot0562} = 112$, which sufficiently indicates that 118, and not 59, should be taken as the atomic weight of tin.

SPECIFIC AND ATOMIC HEATS.

ELEMENT.	Specific Heat.	Atomic Weight.	Atomic Heat.
Bromine (solid)	0·0843	80·	6·7440
Iodine	0·0541	127·	6·8707
Potassium	0·1696	39·	6·6144
Sodium	0·2934	23·	6·7482
Silver	0·0570	108·	6·1560
Gold	0·0324	196·	6·3504
Phosphorus	0·1887	31·	5·8497
Arsenic	0·0814	75·	6·1050
Antimony	0·0508	122·	6·1976
Bismuth	0·0308	210·	6·4680
Sulphur	0·1776	32·	5·6832
Magnesium	0·2499	24·	5·9976
Zinc	0·0955	65·	6·2075
Mercury (solid)	0·0319	200·	6·3800
Copper	0·0951	63·5	6·0389
Lead	0·0314	207·	6·4998
Iron	0·1138	56·	6·3728
Aluminium	0·2143	27·5	5·8932
Tin	0·0562	118·	6·6316
Platinum	0·0324	197·	6·3828

4. *Isomorphism.*—It was discovered by Mitscherlich, that compounds which might be supposed to have a similar chemical composition, very commonly crystallized in the same form, even though quite different in properties. Similar compounds of calcium, strontium, barium, and lead; similar sulphates, selenates, and chromates, and similar phosphates and arsenates, are examples of this law, and many others might be quoted. Compounds which exhibit this peculiarity are said to be *isomorphous* with one another (from ἴσος, like, and μορφή, form). Mitscherlich's law, though not universal, is yet so frequently true that when two analogous substances yield crystals having the same angular measurement, there is a great *a priori* probability that their constituents are arranged in a similar manner, and this probability is sometimes of great use in fixing the atomic weight of an element. For example, the metal aluminium has only one oxide, alumina, which contains $18\cdot\dot{3}$ of the metal to 16 of oxygen. As 16 is the atomic weight of oxygen, it would be natural to suppose that that of aluminium was $18\cdot\dot{3}$, in which case the formula of alumina would be AlO. But alumina is isomorphous with red oxide of iron (ferric oxide) which is known to have the formula Fe_2O_3, and it is therefore concluded that the formula of alumina must be Al_2O_3, in which case the atomic weight of aluminium must be 27·5 Of course either formula agrees equally well with the result deduced from analysis. For if:—

$Al = 18\cdot\dot{3}$, then $AlO = 18\cdot\dot{3}$ Aluminium : 16 Oxygen.
$Al = 27\cdot5$, then $Al_2O_3 = 55$ „ : 48 „

the proportion being the same in both cases. The number 27·5 is confirmed by the specific heat of aluminium, for—

$$\frac{6\cdot3}{\cdot2143} = 29\cdot4$$

which is as near as could be expected.

The atomic weights of the elements, determined by the collation of all these various facts and theories, are given in the table at the commencement of this volume.

EQUIVALENT VALUE OF ATOMS.—ATOMICITY, OR QUANTIVALENCE.

The above remarks will have made it evident that the quantity which is known as the atom of an element is not always the same as its equivalent. The equivalent of oxygen is 8, because 8 parts of oxygen will replace 35·5 of chlorine, or 1 of hydrogen, but the atomic weight of oxygen is held to be 16: the equivalent of nitrogen is 4·6, but its atomic weight is 14; and lastly, in its simplest compound, the equivalent of carbon is 3, whereas its atomic weight is 12. It follows from this, that an atom of oxygen is equivalent to two, an atom of nitrogen to three, and an atom of carbon to four atoms of hydrogen. For if 8 of oxygen be equivalent to 1 of hydrogen, 16 must evidently be equivalent to 2. The atom of each element has therefore its own equivalent value, represented by the number of atoms of hydrogen to which it is equivalent. This equivalent value, as compared with that of the atom of hydrogen, is often spoken of as the *atomicity*, or *quantivalence* of the element, and its amount is described by the words *monad*, *diad*, *triad*, *tetrad*, *pentad*, &c., according as the atom is equivalent to one, two, three, four, or five atoms of hydrogen.

Classification of Elements by their Atomicity.—All those elements of which one atom is equivalent to one atom of hydrogen, either in replacing it or combining with it, are classed together as *monads*. The most important monad elements are, hydrogen, chlorine, bromine, iodine, and fluorine among the non-metals, and potassium, sodium, and silver, among the metals.

The chief *diad* elements are oxygen and sulphur among the non-metals, and barium, strontium, calcium, magnesium, zinc, copper, mercury, and lead among the metals.

As examples of *triads* may be mentioned nitrogen, phosphorus, arsenic, and boron, and the metals bismuth and gold.

Carbon, silicon, and most of the other metals are *tetrads*.

Mode of Marking Atomicity.—As the symbol denotes one atom, it is easy to mark the equivalent value, or atomicity, of that atom by dashes or Roman figures placed above and to the right of the symbol, in the following manner:—H' O'' N'''

C'''' or C^{iv}. The monad atoms, however, are seldom marked at all.

Let it be understood distinctly that a diad atom will take the place of or will combine with any two monad atoms. One atom, or 16 parts of oxygen, for instance, may very often, by mere substitution, be made to *displace* and take the place of two atoms (2 parts) of hydrogen, and we already know that 16 parts of oxygen will *combine with* 2 of hydrogen. The formulæ for a few important compounds will illustrate the different values of the atoms more clearly than words.

HCl, H_2O'', H_3N''', H_4C^{iv}, KCl, K_2O'', $Zn''Cl_2$, $Bi'''Cl_3$.

Mode of Fixing Atomicity.—When an element combines with hydrogen, the hydrogen compound decides the atomicity, as with chlorine, oxygen, nitrogen, and carbon. In other cases the chlorine compound, or any other compound that the element may form with a monad, may be used instead. Thus platinum is reckoned a tetrad because it forms the chloride, $Pt^{iv}Cl_4$, in which the atom of platinum is united with four of the monad atoms of chlorine.

Perissiad and Artiad Atoms.—Those elements whose atomicity is an uneven number are called *perissiads* (περισσός, odd). and those in which it is even, *artiads* (ἄρτιος, even). There is great convenience in the use of these words, which were suggested by Dr. Odling.

Variations of Atomicity.—Many elements have different atomicities in different compounds. For example, a second compound of platinum and chlorine is known, which has the formula $Pt''Cl_2$, and here the metal is clearly a diad. So carbon, which is tetrad in $C^{iv}H_4$ and $C^{iv}O''_2$ (two diad atoms act, of course, like four monad ones), is diad in CO. And nitrogen, which is triad in $N'''H_3$, is monad in N'_2O'', and pentad in $N^{v}H'_4Cl'$. But these variations are subject to one rule, which is of almost universal application. *Perissiads are never artiads, nor artiads perissiads.* A pentad, for example, which of course is a perissiad, may sometimes act as a triad, or a monad, but not as a tetrad, or diad. Two of the oxides of nitrogen, NO and NO_2, are almost the only certain exceptions to this important rule, which must be borne carefully in mind in constructing formulæ.

Radicals.—The formula for nitric acid is HNO_3. Now a large number of compounds called *nitrates* are known, which differ from nitric acid only in containing some metal in place of hydrogen. Thus we have sodium nitrate $NaNO_3$, and silver nitrate AgNO . In all these nitrates, however, the common quantity NO is found, and NO_3 is therefore called the *radical* of the nitrates. In the same way the formula for sulphuric acid is H_2SO_4, and the sulphates all contain the radical SO_4. Neither NO_3 nor SO_4 exist in a separate state. It is probably impossible for them to do so. We do not even know that they exist in the compounds, but we do know that their elements are present in the same proportions in a large number of compounds, and that all these compounds are linked together by a certain similarity of properties; and therefore, without troubling ourselves either to affirm or deny their existence as separate entities, it is convenient to assume their presence in certain compounds. We shall meet with many radicals as we go on, many of them very unlike one another in function; but for the present we have to regard them only in one point of view. In the compounds which contain them, *radicals play the part of elements*, and, like elements, every radical has its own proper atomicity. The acids which contain radicals afford good examples of this.

Nitric Acid	HNO_3	is like	Hydrochloric Acid	HCl.
Sulphuric Acid	H_2SO_4	,,	Water	H_2O.
Phosphoric Acid	H_3PO_4	,,	Ammonia	H_3N.

As Cl is united with one of H, so NO_3 is united with one of H. NO_3 is therefore called a *monad radical*. And as SO_4 is united with two of H, it may be compared to oxygen, and is called a *diad radical*. Their atomicity may be marked as that of elements is marked $(NO_3)'$, $(SO_4)''$, and $(PO_4)'''$, and lastly it must be remembered that in all nitrates NO_3 is monad, and in all sulphates SO_4 is diad.

Measure of the Atomicity of Radicals.—The equivalent value of a radical is measured, just as that of an element is, by the number of atoms of hydrogen or other monad element with which it combines. NO_3, for instance, is a monad radical, because in HNO_3 it is combined with one atom of hydrogen. A radical is in fact an incomplete compound, and it is monad, diad, or triad according as it lacks one, two, or three monads to complete it. The following table includes some of the most important radicals of chemistry. Only a few of them have individual names.

MONAD RADICALS.

HO (Hydroxyl) in Hydrates	as in	H_2O.	KHO.	$Ca''(HO)_2$.	$Bi'''(HO)_3$.
H_2N (Amidogen) in Amides	,,	H_3N.	KH_2N.		
CN (Cyanogen) in Cyanides	,,	HCN.	KCN.	$Hg''(CN)_2$.	
ClO in Hypochlorites	,,	$HClO$.	$KClO$.		
ClO_2 in Chlorites . .	,,	$HClO_2$.	$KClO_2$.		
ClO_3 in Chlorates . .	,,	$HClO_3$.	$KClO_3$.		
ClO_4 in Perchlorates .	,,	$HClO_4$.	$KClO_4$.		
NO_2 in Nitrites . .	,,	HNO_2.	KNO_2.		
NO_3 in Nitrates . .	,,	HNO_3.	KNO_3.		
PO_3 in Metaphosphates	,,	HPO_3.	KPO_3.		

NH_4 (Ammonium) in Ammonium Compounds, as in NH_4Cl. NH_4NO_3.
CH_3 (Methyl) in Methyl Compounds. . . ,, CH_3Cl.
C_2H_5 (Ethyl) in Ethyl Compounds . . . ,, C_2H_5Cl.

DIADS.

SO_3 in Sulphites .	as in	H_2SO_3.	K_2SO_3.	$Ca''SO_3$.
SO_4 in Sulphates .	,,	H_2SO_4.	K_2SO_4.	$KHSO_4$.
CO_3 in Carbonates .	,,	K_2CO_3.	$KHCO_3$.	$Ca''CO_3$.
SiO_3 in Metasilicates	,,	H_2SiO_3.	$Mg''SiO_3$.	

CH_2 (Methylene) in Methylene Compounds, as in CH_2Cl_2.

TRIADS.

PO_4 in Phosphates,	as in	H_3PO_4.	K_3PO_4.	KH_2PO_4.	K_2HPO_4.
AsO_4 in Arseniates	,,	H_3AsO_4.	As_3.	AsO_4.	

CH in such compounds as Chloroform, $CHCl_3$.

TETRADS.

SiO_4 in Silicates . .	as in	H_4SiO_4.	Mg''_2SiO_4.	
P_2O_7 in Pyrophosphates	,,	$H_4P_2O_7$.	$Na_4P_2O_7$.	$Mg''_2P_2O_7$.

NOMENCLATURE.

Names are given to chemical compounds, according to rules which, though not perfect, are better than those employed in any other science. Only a few of the simplest need be indicated in this place.

1. Compounds of two elements only are distinguished by

the termination *ide.* Thus the compound of zinc and oxygen is called oxide of zinc, or zinc oxide; zinc and chlorine, zinc chloride; sodium and chlorine, sodium chloride; and so on. This rule is invariable, though for special reasons a different name is more often used for certain simple compounds.

2. When there are two compounds of the same two elements, the lower—that which contains the least oxygen, chlorine, &c.—is distinguished by the termination *ous*, and the higher by the termination *ic.* The lower chloride of tin $SnCl_2$, for instance, is called stannous chloride, and the higher, $SnCl_4$, stannic chloride. The same rule is applied to acids. HNO_2 is called nitrous acid, and HNO_3 nitric acid.

3. The substances called salts which are derived from an acid which ends in *ous* have the termination *ite.* All the salts, for instance, which contain the radical NO_2, and which therefore correspond to nitrous acid, are called nitrites. Salts which correspond to acids ending in *ic*, have the termination *ate.* Nitric acid, for instance, forms nitrates, sulphuric acid, sulphates, and acetic acid, acetates. Almost the only exception to this rule is in the compounds called hydrates, which contain the radical HO, and which, instead of corresponding in composition to an acid, correspond to water.

Other rules will be more conveniently given as instances occur.

ATOMIC AND MOLECULAR HYPOTHESES.

Atoms.—Atomic Weights.—Dalton, to whom we are indebted for the first clear enunciation of the laws of combination by weight, contrived a beautiful hypothesis to account for them. In its simplest form it is known as Dalton's atomic hypothesis. According to this view, matter is composed of ultimate indivisible particles called *atoms* (ἄτομος, indivisible). Every element has its own atom, peculiar in properties and in weight. All compounds are formed by the union together of elementary atoms. The relative weights of the atoms are expressed by their combining weights, and afford a simple reason for those combining weights. Thus (taking the modern numbers), 35·5 parts of chlorine combine with 1 of hydrogen, because every atom of chlorine is attracted by and combines with one atom of hydrogen. In the same way, every one atom of oxygen, weighing 16, combines with two atoms of hydrogen, weighing 2. In fact, the combining weights of the elements show the relative weights of the atoms, hydrogen being taken as unity, and this is the reason why the combining weights are generally called atomic weights. It follows from this that the ultimate particle—the smallest possible quantity—of any compound must contain at least two different elementary atoms. Compounds are groups of atoms of two or more kinds.

Molecules.—Molecular Weights.—The ultimate particles of compounds, consisting, as we have seen, of groups of atoms, are called *molecules* (*moleculus*, small mass). Their weight is, of course, the sum of the weights of their constituent atoms. Thus the molecular weight of hydrochloric acid, HCl, is 36·5, and of water, 18. Reasons have already been given for believing that the ultimate particles of most elements, at any rate in the gaseous state, consist of two or more similar atoms. The view adopted for compounds must therefore be extended to elements, and we must admit that in the gaseous state elements consist of molecules incapable of division, which molecules themselves consist of atoms of similar kind, sometimes of only one (mercury, zinc, and cadmium gases), sometimes of two (hydrogen, oxygen, &c.), and sometimes of more than two (ozone, phosphorus, arsenic, and sulphur gases).

By an extension of this view, the symbols and formulæ which are used to denote two volumes of any gas, elementary or compound, can also be applied with equal accuracy to denote respectively *one atom* and *one molecule* of it. The atom of hydrogen may be denoted by H, the molecule by H_2; the atom of arsenic by As, the molecule (as gas) by As_4; and single molecules of hydrochloric acid, steam, and ammonia may be represented by the formulæ HCl, H_2O, and H_3N. This leads us to another important hypothesis, which of late years has been very generally adopted. It refers only to gases.

Hypothesis of Avogadro and Ampere.—This hypothesis, suggested by Avogadro in 1811, and further developed by Ampere in 1814, may be stated in the following terms:

Equal volumes of different gases, at equal pressure and temperature, contain equal numbers of molecules.

This hypothesis explains two things.

1. The fact already mentioned (pages 34, 56), that all gases are equally affected in volume by variations of pressure and temperature. Physical agencies, such as heat and pressure, do not affect the nature of molecules, but only their distance from one another, and this being the same in all gases under similar conditions, the effect produced by those agencies must naturally be the same in all cases.

2. The law of gaseous volumes. If one molecule of chlorine enters into chemical action with one molecule of hydrogen, the same will be true of one thousand, or one million molecules of each. Now by the hypothesis one thousand, or one million, molecules of chlorine occupy a space equal to that occupied by one thousand, or one million molecules of hydrogen, and it therefore follows that the two elements should react together in the proportion of equal volumes, which experiment proves to be the fact. In like manner, two volumes of hydrogen react with one volume of oxygen, because one volume of oxygen contains *n*, and two volumes of hydrogen 2*n* molecules, and every molecule of oxygen reacts with two molecules of hydrogen.

Structure of Gases.—Summary of Hypotheses.—The modern doctrine may be summed up in these words.

1. *Equal volumes of gases, at equal pressure and temperature, contain equal numbers of molecules.*

2. *Every molecule consists of one or more atoms, either similar in kind (elementary molecules), or dissimilar in kind (compound molecules).*[*]

---

[*] The molecular hypothesis is sometimes referred to so loosely, that a few further remarks upon it may not be out of place.

1. The hypothesis asserts nothing of the relative bulk of molecules, or atoms. It does not assert that one molecule of hydrogen has the same bulk as one molecule of chlorine, but only that *n* molecules of hydrogen occupy, by their molecular movements, the same space as *n* molecules of chlorine.

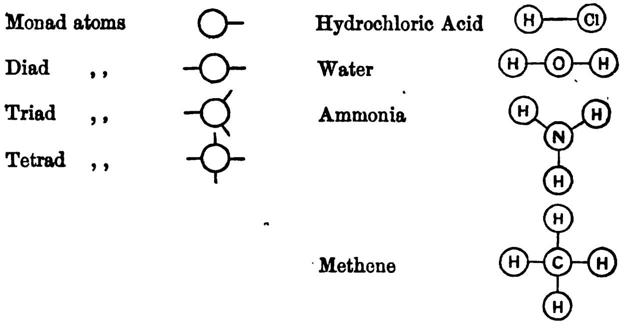
2. Similarly, the hypothesis does not assert that the bulk of an atom of hydrogen is equal to the bulk of an atom of chlorine, but only that two atoms of hydrogen and two atoms of chlorine occupy, respectively, by their atomic movements, the space of one molecule. The atoms of matter may be almostly infinitely small, as compared with the molecules, and if the bulk of the molecules were known we should be no nearer to knowing that of the atoms.

A molecule of hydrogen may, for aught we know, be comparable to an inflated bladder, with two peas rotating in it.

Molecules of Solids and Liquids.—The formulæ applied to gases are generally extended for convenience sake, but by imperfect reasoning, to the same substances in the liquid and solid state. But it is right to point out that neither hypothesis nor experiment tell us anything of the number of atoms contained in the molecules of solids and liquids. We have good reason for believing that a molecule of iodine vapour contains only two atoms; but, for aught we know, a molecule of solid iodine may contain a thousand. The formula for common salt, which is non-volatile, is written NaCl, on account of its analogy, not with solid, but with *gaseous* HCl. The formula NaCl does indeed tell accurately, as far as our knowledge goes, the *relative* number of atoms of sodium and chlorine that are present in the compound, but it does not tell us the *absolute* number that are present in each molecule. The true formula for common salt, as compared with that of hydrochloric acid gas, HCl, would be $Na_n Cl_n$, n standing for an unknown and probably very large number. In other words, the formula for common salt, or for any other solid or liquid, is probably only the formula for the nth part of one molecule. There is both convenience and propriety in the use of the ordinary formulæ for solids and liquids, inasmuch as they truly represent the relative quantities of matter which are concerned in chemical changes; but it should not be forgotten that these formulæ differ from those of gases in that the latter tell us, *in addition*, how much matter is present in a certain space, and also, if the molecular hypothesis be adopted, the absolute number of atoms in each molecule.

Hypothesis to Account for Atomicity.—The different equivalent value of different elementary atoms can be accounted for by supposing that each atom has the power of fixing to itself a certain number of other atoms. It is as though a carbon atom, for example, had four arms by which it could grasp and be grasped by four hydrogen, or chlorine atoms, or two oxygen, or sulphur atoms, each of the latter having two arms of its own. It must, of course, be understood that this is a mere illustration, for no one believes that the atoms have real arms projecting from them. But the illustration gives a lively idea of the hypothesis; and atoms are sometimes figured as circles with arms (called bonds) proceeding from them.

The following figures represent a few of these atoms and their compounds:



The variations of atomicity by pairs, which have been described before (page 85), are explained by supposing that the bonds of an atom have the power, under certain circumstances, of neutralizing one another in pairs. To carry out the simile used before, the carbon atom has four arms, and is therefore a tetrad; but if two of those arms are clasped together, they will not so easily grasp other atoms, and the carbon atom will therefore act as a diad.

The hypothesis, put in this form, appears somewhat fanciful, but it harmonizes well with known facts, and has the merit of assisting wonderfully in the comprehension of the complex compounds of organic chemistry.

EQUATIONS, OR FORMULÆ OF CHEMICAL CHANGE.

Chemical changes of all kinds can be very conveniently represented in the form of equations, the formulæ for the substances concerned being written in the first half, and the formulæ for the new substances produced in the other. Thus the combination of zinc and chlorine is thus expressed:

$$\underset{\text{Zinc.}}{Zn''} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\text{Zinc Chloride.}}{Zn''\ Cl_2};$$

which may be read in the following manner. One molecule of zinc added to one molecule (two atoms) of chlorine is equal to, or rather produces one molecule of zinc chloride.

Such equations are often called *formulæ*, which introduces a little confusion, since we have already seen that the term formula is also applied to the aggregate of symbols which denotes the molecule of a single element or compound. In making out an equation of this kind, it is necessary, as with algebraical equations, to take care that every atom which appears in the first half shall duly appear and be accounted for in the second. The following equation, which expresses the oxidation of ethylene gas, illustrates this:

$$\overset{\text{Ethylene.}}{C_2H_4} + 3\,O_2 = \overset{\text{Carbonic Anhydride.}}{2\,CO_2} + \overset{\text{Water.}}{2\,H_2O}.$$

Two atoms of carbon, four of hydrogen, and six of oxygen are concerned in the change, and it will be seen that though differently arranged, they are all accounted for in the second half of the equation. The large figures in the above and in similar equations refer, it must be remembered, to the *whole molecule*. $2CO_2$, for instance, means two molecules of carbonic anhydride, containing together two atoms of carbon and four of oxygen. A few more equational formulæ, one or two of them rather complex ones, may, with advantage, be studied in this place. They represent chemical changes of several kinds.

Decomposition of water:

$$2\,H_2O'' = 2\,H_2 + O_2.$$

Synthesis of water:

$$2\,H_2 + O_2 = 2\,H_2O.$$

Substitutions:

$$\overset{\text{Sulphuric Acid.}}{H_2SO_4} + Zn'' = \overset{\text{Zinc Sulphate.}}{Zn''SO_4} + H_2.$$

$$\overset{\text{Water.}}{2\,H_2O''} + K_2 = \overset{\text{Potassium Hydrate.}}{2\,KHO''} + H_2.$$

$$\overset{\text{Copper Chloride.}}{Cu''Cl_2} + Fe'' = \overset{\text{Ferrous Chloride.}}{Fe''Cl_2} + Cu''.$$

Double decompositions:

$$\overset{\text{Mercuric Chloride.}}{Hg''Cl_2} + \overset{\text{Potassium Iodide.}}{2\,KI} = \overset{\text{Mercuric Iodide.}}{Hg''I_2} + \overset{\text{Potassium Chloride.}}{2\,KCl}.$$

Double decompositions—*continued.*

$$\overset{\text{Potassium Hydrate.}}{KHO''} + \overset{\text{Hydrochloric Acid.}}{HCl} = \overset{\text{Water.}}{H_2O''} + \overset{\text{Potassium Chloride.}}{KCl.}$$

$$\overset{\text{Copper Oxide.}}{Cu''O''} + \overset{\text{Nitric Acid.}}{2\,HNO_3} = \overset{\text{Water.}}{H_2O''} + \overset{\text{Copper Nitrate.}}{Cu''(NO_3)_2.}$$

$$\overset{\text{Calcium Chloride.}}{3\,Ca''Cl_2} + \overset{\text{Phosphoric Acid.}}{2\,H_3PO_4} = \overset{\text{Hydrochloric Acid.}}{6\,HCl} + \overset{\text{Calcium Phosphate.}}{Ca''_3(PO_4)'''_2.}$$

The brackets in the two last examples will readily be understood. $Cu''(NO_3)_2$ expresses that one atom of the diad metal copper is united with two units of the monad radical NO_3.

One great advantage of these equations is that they afford us the means of calculating the respective quantities by weight in which bodies act on one another. Take, for example, the above-described action of mercuric chloride and potassium iodide. $HgCl_2$ means one atom of mercury weighing 200 combined with two atoms of chlorine weighing $35{\cdot}5 \times 2 = 71$, total 271. KI means one atom of potassium 39 and one atom of iodine 127, total 166; or, as 2KI is employed, 332. We therefore know that 271 parts (pounds, grammes, or tons) of mercuric chloride will act upon 332 parts of potassium iodide, and in a similar manner we can calculate that 454 parts of mercuric iodide (HgI_2 = mercury, 200 + iodine, $127 \times 2 = 254$), and 149 parts of potassium chloride will be produced during the change. Now suppose that a chemist has 100 grains of mercuric chloride, and wishes to know how much potassium iodide he must add to convert all the mercury into iodide. He knows that 271 parts of the chloride will require 332 of the iodide, and he has therefore only to perform a simple proportion sum.

$$271 : 332 :: 100 : x = 122\tfrac{1}{2} \text{ grains.}$$

In like manner he can readily find out that he ought to obtain as the result of the action $167\frac{1}{2}$ grains of mercuric iodide and 55 grains of potassium chloride, for as 271 parts of mercuric chloride yield 454 parts of mercuric iodide and 149 parts of potassium iodide,

$$271 : 454 :: 100 : x = 167\tfrac{1}{2} \text{ grains}$$
$$\text{and } 271 : 149 :: 100 : x = 55 \quad \text{,,}$$

In short, to sum up the whole reaction, 100 grains of mercuric chloride, with 122½ grains of potassium iodide, will yield 167½ grains of mercuric iodide and 55 grains of potassium chloride. It is of course equally easy to find out how much mercuric chloride and potassium iodide must be employed to yield a certain weight, say 100 grains of mercuric iodide.

When the equations refer to gases they have another advantage. As every single formula for an element or compound denotes two volumes of gas, the volumes of different gases concerned in a reaction can at once be inferred from the equation. In the equation given above

$$C_2H_4 + 3\,O_2 = 2\,CO_2 + 2\,H_2O,$$

C_2H_4 means 2 volumes of ethylene, and $3O_2$ means $3 \times 2 = 6$ volumes of oxygen; so that 2 volumes of ethylene require 6 volumes of oxygen for their complete oxidation. There will be produced during the action, $2CO_2$, that is, 4 volumes of carbonic anhydride and $2H_2O$, that is, 4 volumes of steam. For if CO_2 represents 2 volumes, $2CO_2$ must evidently represent 4 volumes. It must not, however, be forgotten that these relations of volume are only correct if the pressure and temperature remain unaltered. If otherwise, a correction must be made by the methods already given (pages 34, 57).

ACIDS, BASES, SALTS.

Allusion has already been made (pages 49, 50) to certain substances which have long been known under the respective names of *acids* and *bases*. Acids are sour and redden litmus, and bases, when soluble, have what is called an alkaline taste, and turn the colour of reddened litmus back again to blue. They have a kind of antagonistic function, and will *neutralize* one another. When an acid acts on a base, a new compound called a *salt* is produced, which commonly has no action on either blue or red litmus.

But although in well-marked cases acids, bases, and salts are so different from one another in properties, modern chemistry has taught us that the compounds usually known by those names bear so much resemblance to one another and to other compounds in *structure*, that it is impossible to frame accurate definitions for them. If possible it would perhaps be as well to banish the terms altogether; but as that would produce great inconvenience, it is better to take such imperfect definitions as we can get.

1. *Acids.—An acid is composed of hydrogen with one of those radicals, elementary or compound* (page 86), *which are called acid radicals. The hydrogen can be replaced by metals, in which case one of the compounds called salts is formed. Acids redden litmus, and are commonly sour.*

The following are important examples:—Hydrochloric Acid, HCl; Nitric Acid, HNO_3; Sulphuric Acid, H_2SO_4; Phosphoric Acid, H_3PO_4. In the first of these hydrogen is united with the elementary acid radical Cl, in the others, with the compound radicals NO_3, SO_4, and PO_4.

The imperfection of the definition will be apparent if we remember that we can only define an acid radical as a radical which, united with hydrogen, forms an acid.

Basity of acids.—Acids are said to be *monobasic*, *dibasic*, or *tribasic*, according as they contain one, two, or three atoms of hydrogen which can be replaced by acids. In the previous examples HCl and HNO_3 are monobasic, H_2SO_4 is dibasic, and H_3PO_4 is tribasic. Of course this is equivalent to saying that the radicals of those acids are monad, diad, or triad. PO_4 is a triad radical, and its acid may be written in this way, $H_3(PO_4)'''$. Acids which contain more than one atom of replacable hydrogen are said to be *polybasic*.

Oxygen Acids, or Oxy-acids.—Most acids contain oxygen as a part of their radical, and these are related in a very simple manner to a particular series of oxides, which, for that reason, are called *acid oxides* or *anhydrides*. When one of these anhydrides combines with water, an acid is formed, and when, on the other hand, water is removed from an acid, the corresponding anhydride is obtained.

$$\overset{\text{Nitric Anhydride.}}{N_2O_5} + H_2O = \overset{\text{Nitric Acid.}}{2\,HNO_3}.$$

$$\overset{\text{Sulphuric Anhydride.}}{SO_3} + H_2O = \overset{\text{Sulphuric Acid.}}{H_2SO_4}.$$

$$\overset{\text{Phosphoric Anhydride.}}{P_2O_5} + 3\,H_2O = \overset{\text{Phosphoric Acid.}}{2\,H_3PO_4}.$$

One molecule of nitric anhydride, uniting with one of water, forms two molecules of nitric acid. In a similar manner, if from one molecule of sulphuric acid one molecule of water be taken, one molecule of sulphuric anhydride will remain.

It must not, however, be supposed the acids are always, or even generally, prepared in practice from the anhydrides. One anhydride, indeed, silicic anhydride, SiO_2, will not combine directly with water, although its acid can be obtained by indirect means. The anhydrides of many acids have not yet been obtained, and one anhydride (carbonic anydride, CO_2) is known to which no corresponding acid can be proved to exist. The formula for the acid should be H_2CO_3.

2. *Salts.—A salt is a compound containing a metal and an elementary or compound acid radical.*

A salt only differs from an acid by containing a metal in place of hydrogen. Taking, for example, the salts of sodium, those corresponding to the acids already mentioned are—

$Na\,Cl$	Sodium	chloride;	corresponding to	$H\,Cl$.
$Na\,N\,O_3$	,,	nitrate	,,	$H\,N\,O_3$.
$Na_2\,S\,O_4$	,,	sulphate	,,	$H_2\,S\,O_4$.
$Na_3\,P\,O_4$	,,	phosphate	,,	$H_3\,P\,O_4$.

The proportion which the metal and radical bear to one another depends of course on the atomicity of each. Thus the chloride, nitrate, sulphate, and phosphate of the diad metal calcium, and of the triad metal bismuth, are formulated in this way:

$Ca''\,Cl_2$	$Ca''\,(N\,O_3)_2$	$Ca''\,(S\,O_4)''$	$Ca''_3\,(P\,O_4)'''_2$.
$Bi'''\,Cl_3$	$Bi'''\,(N\,O_3)_3$	$Bi'''_2\,(S\,O_4)''_3$	$Bi'''\,(P\,O_4)'''$.

In many salts of polybasic acids the hydrogen is only partially replaced by metals. Thus we have

Phosphoric Acid.	Sodium-dihydrogen Phosphate.	Disodium-hydrogen Phosphate.	Sodium Phosphate, or Trisodium Phosphate
$H_3\,P\,O_4$.	$Na\,H_2\,P\,O_4$.	$Na_2\,H\,P\,O_4$.	$Na_3\,P\,O_4$.

It will be seen from the above remarks that acids are really hydrogen salts. Some chemists indeed name them accordingly, and call H_2SO_4 "hydrogen sulphate," instead of sulphuric acid. The term acid is applied by them to the anhydrides, so that the formula for sulphuric acid becomes SO_3. Many salts are known which are more or less irregular in their composition, but modern chemical theory enables us to give a tolerably satisfactory account of all of them.

3. *Bases.—A base is a metallic hydrate* (that is, a compound of a metal with the radical HO) *which is capable of reacting with acids to form salts.*

As HO is a monad radical, the constitution of hydrates is closely analogous to that of chlorides, nitrates, &c. For example:

Sodium Hydrate.
$NaHO$ is like $NaCl$ and $NaNO_3$.

Calcium Hydrate.
$Ca''(HO)_2$ is like $Ca''Cl_2$ and $Ca''(NO_3)_2$.

Bismuth Hydrate.
$Bi'''(HO)_3$ is like $Bi'''Cl_3$ and $Bi(NO_3)_3$.

Bases, like acids, are related to a particular series of oxides called *basic oxides.* These oxides are sometimes called bases. They differ from the true bases by the elements of water.

Sodium Oxide.				Sodium Hydrate.
Na_2O	$+$	H_2O	$=$	$2\,NaHO$.
Calcium Oxide.				Calcium Hydrate.
$Ca''O$	$+$	H_2O	$=$	$Ca''(HO)_2$.
Bismuth Oxide.				Bismuth Hydrate.
$Bi'''_2O''_3$	$+$	$3\,H_2O$	$=$	$2\,Bi'''(HO)_3$.

Some basic oxides, however (bismuth oxide, for instance), will not combine directly with water; and, on the other hand, some basic hydrates (sodium hydrate, for instance) cannot be dehydrated by heat. Calcium oxide and hydrate are examples of compounds which experience both changes with ease. The oxide (quick lime) combines eagerly with water, great heat is produced, and calcium hydrate (slacked lime) is formed. When slacked lime is heated, water is expelled, and quick lime once more obtained.

$$Ca''O + H_2O = Ca''(HO)_2;$$
$$\text{and}\quad Ca''(HO)_2 - H_2O = Ca''O.$$

Formation of Salts.—Salts can be formed by a variety of processes, only a few of which can be specified here, with one or two examples of each. It must not be supposed that every process is practicable in all cases.

1. By the action of metals on acids:

$$H_2SO_4 + Zn'' = Zn''SO_4 + H_2.$$

2. By the action of metallic oxides on acids:

$$Ca''O + 2\,HCl = Ca''Cl_2 + H_2O.$$

3. By the action of bases on acids:

$$NaHO + HCl = NaCl + H_2O;$$
$$Ca''(HO)_2 + 2\,HCl = Ca''Cl_2 + 2\,H_2O.$$

This is the case before referred to.

4. By the action of anhydrides on metallic oxides:

$$SO_3 + Na_2O = Na_2SO_4.$$

5. By the action of anhydrides on bases:

$$CO_2 + 2\,NaHO = Na_2CO_3 + H_2O.$$

It will be seen that in most of these reactions water is formed simultaneously with the salt.

PART II.

NON-METALLIC ELEMENTS.

INTRODUCTION.

THE old distinction between organic and inorganic chemistry is fast fading away. It was formerly believed that vegetables and animals had the power of producing in their organisms chemical compounds which could not be formed artificially in the laboratory. Vast numbers of compounds were known which owed their origin directly or indirectly to the animal or vegetable kingdom, and which could not be obtained from any other source. To the department of chemistry which dealt with such compounds, the name "Organic Chemistry" was very properly applied. But the progress of scientific research has taught us to manufacture a great number of these compounds from inorganic materials, and there is therefore no longer any reason why they should be separated from other compounds in a general system of classification.

The so-called organic compounds have, however, one feature in common. They all contain carbon, and it is therefore convenient for purposes of study to retain them in a separate department of chemistry. Moreover, although some few carbon compounds enter into the composition of important minerals, a very large number of them bear some direct or indirect relation to the processes of life, and the term "organic" may therefore, in this limited sense, still be applied to them. It is only necessary to bear in mind that what is still generally called organic chemistry is but the chemistry of carbon compounds. In the present division of the book we shall say as little about it as possible.

A less important distinction is that which is commonly made between metallic and non-metallic elements. No exact definition can be attached to the word metal, no one property can be connected exclusively with it, and there is no line of demarcation that can be drawn between the two classes. We shall trouble ourselves very little with the precise meaning of the word metal, but shall use it in its current sense, and shall introduce the elements in the order which appears most simple and convenient, grouping them according to their atomicities (page 84). The present part contains an account of the elements generally classed as non-metallic. Part III. is devoted to the metals, and Part IV. to organic chemistry.

CHAPTER I.

NON-METALLIC MONADS.

(Hydrogen, Chlorine, Bromine, Iodine, Fluorine.)

HYDROGEN.

Symbol, $H = 1$. Formula, H_2.

THIS element in its chemical relations resembles the monad metals. It stands alone among the non-metals, and it must, for convenience, be studied first.

PREPARATION.

Hydrogen occurs in nature almost exclusively in a state of combination, but various means are known by which it can be set free and obtained in a state of purity. We have already seen that when a current of electricity from a tolerably powerful battery is transmitted through water, slightly acidulated with sulphuric acid, the water is decomposed, oxygen goes to one pole of the battery and hydrogen to the other; and both gases can easily be collected (p. 12). Many other compounds containing hydrogen will yield it up under similar treatment.

The ordinary processes for preparing hydrogen almost all involve the use of some metal. A certain number of the metals have the power of displacing hydrogen from its combinations, either easily or with difficulty. For common purposes zinc or iron is used.

Experiment 1.—Boil some water for fifteen minutes, that all the air contained in it may be expelled; let it cool, and fill a bowl and a test-tube with it; close the latter with the finger, and invert the mouth of it under the water in the bowl. Now fasten to a wire a piece of sodium, of the size of a small pea, and thrust it quickly under the mouth of the test-tube; the metal frees itself from the wire, and as it is lighter than water, it ascends into the tube, floating there with a rotatory motion; a gas is evolved from the water, and collects in the upper part of the tube. This gas is hydrogen. The metal sodium displaces one half of the hydrogen from water, in the manner shown in the following formula:

$$2\ H_2O + Na_2 = 2\ NaHO + H_2.$$

The compound NaHO is called sodium hydrate. It remains dissolved in the water, and may be obtained in a pure state if the water is evaporated off. Close the tube again with the finger, remove it from the bowl, and apply a light to the mouth, the gas will burn with a flash of light. This experiment proves hydrogen to be a *combustible* gas. Pour into the bowl some solution of litmus, which has been reddened with a drop of vinegar or other acid, the litmus will be immediately changed to blue, showing that sodium hydrate is one of the substances called *bases* (pp. 50, 98).

Experiment 2.—Lay a piece of blotting-paper on the surface of some water contained in a saucer, and throw upon it a small piece of sodium, an energetic decomposition of the water will take place, and in a few moments the sodium will apparently burst into flame, and burn for some time with an intense yellow colour. This apparent combustion of the sodium is really due to the burning of the hydrogen, set free by that metal, which is inflamed by the intense heat which accompanies its evolution. This experiment differs only from the preceding one inasmuch as in the former case the hydrogen is collected, while in the latter it is burnt as it is liberated. The sodium hydrate may be rendered evident as before by the addition of red litmus solution to the water.

Experiment 3.—Throw a piece of potassium on some water (the blotting-paper may be dispensed with in this case), the same effect as in the preceding experiment will occur, the water will be more violently decomposed, and the hydrogen —but apparently the potassium—will burn instantaneously with a beautiful violet flame. Potassium hydrate, KHO, remains in solution in the water. It is a base like sodium hydrate.

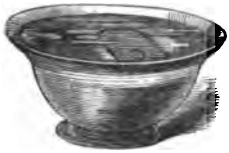
Experiment 4.—What sodium and potassium accomplish at ordinary temperatures, iron can do if it be heated to redness. Pass water in the form of steam, obtained by boiling the water in a flask, or a retort, through a red-hot iron pipe, as a gun or gas barrel, containing small iron nails or wire. At this high temperature the iron in the pipe unites with the oxygen in the water, forming a black oxide of iron, and the hydrogen is set free and may be collected in the manner described below. The reaction is as follows:

$$4\ H_2O + 3\ Fe = Fe_3O_4 + 4\ H_2.$$

The best means, however, for obtaining hydrogen is by the decomposition of an acid by iron or some other metal, zinc being generally chosen for the purpose.

Collection of Gases.—For collecting considerable quantities of gases the following contrivance, called a *pneumatic trough*, may be used. Make a shelf out of slate or a piece of lead, from three to four inches broad, and so long that it will rest about half way up the sloping sides of a pan or wash-hand basin; cut a hole about half an inch in diameter through the centre of the shelf, and having placed the latter in position, pour into the pan sufficient water to cover it an inch deep. The shelf is for the purpose of supporting the vessel intended for the reception of the gas, which, filled with water, is placed with its mouth exactly over the hole. The gas to be collected is then delivered from a tube, the extremity of which is placed directly under the mouth of the inverted vessel. The accompanying figure conveys an idea of the kind of apparatus required.

Fig. 38.



Experiment 5.—Put half an ounce of *granulated* zinc * in a bottle or flask, and pour over it a little water. No action takes place, but if a small quantity of sulphuric acid be gradually added, effervescence, and heating of the mixture will ensue. The effervescence is caused by the escape of

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* Zinc is granulated by being melted, in an iron ladle or spoon, and poured from the height of a few feet into cold water.

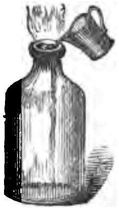
hydrogen gas. Insert into the mouth of the bottle a cork, which has previously been perforated and fitted with a tube bent so that the mouth of it may be conveniently placed beneath the hole in the shelf of the pneumatic trough.* Allow time for the hydrogen to displace the air contained in the flask and bent tube (two or three minutes is sufficient if the effervescence is tolerably brisk). While the air is being expelled, preparation may be made for collecting the gas by filling several bottles quite full of water. One of the bottles is then closed with a smooth card or glass plate, and rapidly inverted in the pneumatic trough with its mouth directly over the hole in the shelf. The tube being then placed beneath, the gas will ascend and displace the water in the bottle. When the entire displacement is effected the bottle is corked or stoppered while still in the trough, removed and replaced by another bottle, inverted in the same manner, and so on until the evolution of gas ceases.

There is one indispensable caution to be observed in experimenting with hydrogen, which is, *not to begin to collect the gas until all the atmospheric air existing in the flask has been expelled*, as otherwise an explosion might take place. To be quite safe, it is as well to reject the first bottleful of hydrogen collected.

PROPERTIES.

Experiment 6.—Inflame hydrogen contained in a bottle, and immediately pour in some water. The water does not extinguish the flame, but rather increases it, since it rapidly forces the gas out of the flask. The gas does not burn in the interior of the vessel, but only on the outside, where it is surrounded by atmospheric air.

Fig. 39.



Experiment 7.—Open a bottle of hydrogen under an inverted tumbler, and, after a minute or two, apply a lighted taper to the mouth of the tumbler. A flame will burst forth from the tumbler with a slight report. The gas has ascended from the bottle into the tumbler, and is consequently *lighter* than common air. In this experiment the lower vessel must not be immediately exposed to the lighted taper, because, if all the hydrogen is not displaced, an explosion might ensue that might break the bottle; but if the taper be applied after ten minutes have elapsed, the bottle will be found no longer to contain any combustible gas, the gas having entirely escaped.

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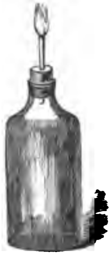
* Instead of a rigid glass tube, it is more convenient for this and similar purposes to carry the gas to the pneumatic trough by means of a piece of india-rubber tubing, of about one-eighth of an inch internal diameter. It may be slipped over a short piece of glass tubing which passes through the cork.

Hydrogen is the *lightest* of all gases. Its specific gravity is 1, and 14½ measures of it weigh only as much as one measure of atmospheric air. On account of this lightness, it may be used for filling balloons.

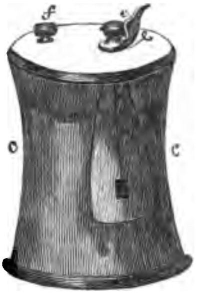
Fig. 40.



Experiment 8.—If, instead of the glass tube, a piece of tobacco-pipe be adapted to the cork of the flask from which hydrogen was evolved, and the gas then lighted, it will burn like a taper. To kindle the gas, instead of a match or a taper, very finely divided platinum may be employed. This can be prepared in a few minutes by dropping a solution of platinum chloride on blotting-paper, attaching it to a wire, and igniting it over a spirit-lamp, till nothing but a grey coherent ash remains. The platinum is thus reduced to an extremely minute state of subdivision, and in this state it exhibits the remarkable property of igniting in hydrogen and inflaming it. It is called spongy platinum, and is employed as tinder in the well-known *Döhbereiner's lamp.*

The apparatus here represented consists of a flask, having the bottom broken off, and to the neck of which the cover of the glass vessel, *c*, with the cock, *e*, is fastened air-tight. A piece of zinc is suspended in the flask by means of a wire. If diluted sulphuric acid is now poured into the vessel, *c*, upon which the cover with the flask attached is placed, then, the cock being opened, that the air contained in the flask may be displaced by the acid from beneath, hydrogen is immediately evolved by the contact of the zinc with the acid, which hydrogen must be collected in the flask by closing the cock, *e*, the acid being thereby forced into the exterior vessel, until it no longer touches the zinc. Upon opening the stop-cock, *e*, the gas issues from the fine

Fig. 41.



jet, and is directed against the spongy platinum, *f*. As the gas escapes, the sulphuric acid passes again into the interior vessel, and generates fresh hydrogen upon reaching the zinc. Spongy platinum possesses, in a high degree, the power of absorbing oxygen and condensing it within its pores; if hydrogen be then presented to it, these two gases will be brought into such intimate contact, by the powerful force of attraction, that they will chemically combine to form water, and the heat thus liberated is sufficient to ignite the platinum tinder, and to inflame the gas, which subsequently issues from the jet. Many aeriform bodies, which do not freely unite with each other, can be forced to combine by means of spongy platinum.

Experiment 9.—To observe the remarkable lightness of hydrogen, a small balloon of gold-beater's skin may be filled with hydrogen by means of the apparatus used in Experiment 5. The balloon is squeezed flat to expel the air, and the tube delivering the gas is passed a short way into its orifice, and secured there with a piece of thread. When the inflation is complete, the tube is withdrawn, and the twine tightened simultaneously, thus preventing any escape of gas. The balloon, if unimpeded, will then ascend to a great height. It may be made captive with a piece of thread.

Experiment 10.—Pour the contents of the flask in which the hydrogen was generated (Experiment 5) into a porcelain dish, boil until they are reduced in bulk to one-half or thereabouts, and filter them (page 49). A black residue will remain on the filter, which consists of the impurities contained in the zinc; the zinc itself has been dissolved, and has been converted into a *salt* (page 97), called zinc sulphate, which, on the cooling of the solution, is deposited in colourless crystals. The reaction which takes place between zinc and sulphuric acid is represented by the following equation:

$$\underset{\text{Sulphuric Acid.}}{H_2(SO_4)} + \underset{\text{Zinc.}}{Zn} = \underset{\text{Zinc Sulphate.}}{Zn(SO_4)} + \underset{\text{Hydrogen.}}{H_2}.$$

CHLORINE.

Symbol, Cl = 35·5. Formula, Cl_2.

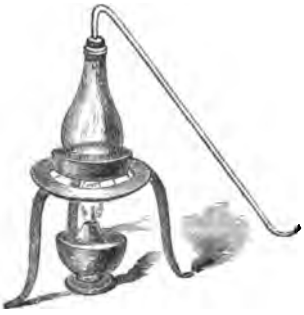
Chlorine is one of a group, the members of which are characterized by a remarkable similarity of chemical properties. It consists of the elements chlorine, bromine, iodine, and fluorine, the three first of which are often termed *halogens* (from ἅλς, sea-salt), in allusion to their marine origin.

Chlorine is a greenish-yellow gas, 2½ times heavier than air, suffocating and irrespirable unless very much diluted with air. Its odour, when very dilute, is somewhat like that of sea-weed, a peculiarity which it shares with the other halogens. Chlorine occurs only in combination; chiefly with sodium as common salt (sodium chloride, NaCl), which forms immense deposits in England and elsewhere, and is the chief ingredient of sea-water.

PREPARATION.

Experiment 1.—Pour one ounce and a half of hydrochloric acid upon a quarter of an ounce of finely-powdered black oxide of manganese, and heat it gradually in a flask, to which is adapted a bent glass tube; a yellowish-green gas is disengaged, which is collected by the process already described. The pneumatic trough is, however, filled with warm water instead of cold. This gas is *chlorine* (from χλωρός, *green*). Fill with it several six-ounce bottles of white glass, and cork them up. Fill, likewise, a bottle with two-thirds of chlorine and one-third of water, and shake it up; suction is exerted upon a finger which closes the mouth of it,—a proof that a vacuum has been produced. If the finger be removed, the air immediately rushes in. This vacuum was caused by the chlorine having dissolved in the water, which might be inferred also from the disappearance of the yellow colour

Fig. 42.



from the upper part of the bottle. One measure of cold water dissolves two measures of chlorine. This solution is called *chlorine water.*

The mode in which chlorine is formed in this experiment is shown in the following formula:

$$\underset{\text{Manganese Peroxide.}}{MnO_2} + 4\,HCl = \underset{\text{Manganese Chloride.}}{Mn\,Cl_2} + 2H_2O + Cl_2.$$

When the evolution of gas has quite ceased the liquid in the flask may be filtered and evaporated, when it will yield on cooling pink crystals of $MnCl_2$.

Experiment 2.—Chlorine may also be prepared from common salt by mixing three quarters of an ounce of it with half an ounce of black oxide of manganese, two ounces of sulphuric acid, and one ounce of water, and heating the mixture:

$$\underset{\text{Sodium Chloride.}}{2\,NaCl} + \underset{\text{Manganese Peroxide.}}{MnO_2} + \underset{\text{Sulphuric Acid.}}{2\,H_2SO_4} = \underset{\text{Sodium Sulphate.}}{Na_2SO_4} + \underset{\text{Manganese Sulphate.}}{MnSO_4} + 2\,H_2O + Cl_2$$

Chlorine acts as a *poison* on being inhaled; hence, *care must be taken not to inhale it* while preparing it. For greater security, pour some drops of alcohol and ammonia upon a cloth and wave it frequently in the air; the chlorine contained in the air will then be so altered that it will lose its injurious properties.

PROPERTIES.

Experiment 3.—In order to recognise the odour of chlorine, smell chlorine water (but not the gas) cautiously; the chlorine water may be tasted also without danger.

Experiment 4.—If a flask containing chlorine gas be exposed to the air for a short time, no diminution of the chlorine will be perceptible; but if the flask be inverted it will soon contain only atmospheric air. Chlorine is two and a half times *heavier* than common air, and may be easily poured from one vessel to another like water without material waste; its specific gravity is 35·5 (H = 1).

Experiment 5.—Introduce a piece of litmus-paper into chlorine gas, and it becomes white; pour chlorine water upon red wine, or ink, and both the liquids will lose their colour. *Chlorine bleaches and destroys most colours derived* *from the animal or vegetable kingdom.* In consequence of this property, chlorine has become a most important agent in bleaching; and linen, cotton, paper, and other materials, may be rendered perfectly white by it in a few hours; while, by the old method of laying them on the grass in the sun, weeks, and even months, were required for effecting it. Substances called *antichlors* are sometimes used to remove the last traces of chlorine. Sodium hyposulphite is the most powerful. The modern method of bleaching is very excellent, and does not in the least injure the strength of the fabric, provided all the chlorine be completely removed after the bleaching is finished, which is not so easily done as might be supposed. If this precaution is not observed, or if the chlorine water is too strong or in excess, then, indeed, after the colour is destroyed, the fibres of the yarn or fabric itself will be attacked. The substance commonly called *chloride of lime* is now used instead of chlorine. It is a salt from which chlorine is easily disengaged, even by mere exposure to the air.

Experiment 6.—Apply chlorine water to decaying and nauseous substances (water in which flowers have been kept, manure, rotten eggs, &c.); the bad odour will at once entirely vanish. Thus it not only *decomposes* colours, but also the *volatile combinations* formed during putrefaction, and which occasion disagreeable odours. It acts in a similar manner also upon morbific matter (matters of contagion, miasmata), which, being diffused in the air or attached to clothes and beds, may communicate disease. Chlorine is, therefore, a powerful *disinfecting agent*, and is used for purifying all putrefying matter and infected atmospheres, and for arresting the decay of organic substances. Musty casks may also be purified by washing them first with chlorine water, and then with some milk of lime. Mouldy cellars, in which milk or beer cannot be kept without turning sour, are again rendered serviceable for a long time by fumigating them with chlorine gas, or by washing them with chlorine water, or a solution of chloride of lime.

Experiment 7.—Fill a small bottle with chlorine water, and invert it in a vessel filled with water; if this is put away in a dark place, it remains unchanged; but if it is exposed to the sun, a colourless gas will collect in the upper part of the flask, in which a glowing taper will inflame; this gas is oxygen. After some days the water will entirely lose its odour of chlorine, and will have acquired a sour taste, and instead of bleaching blue litmus-paper, it will redden it. Three elements only were present, the constituents of water and chlorine; thus it is obvious that the chlorine must have united with the hydrogen of the water to form hydrochloric acid, the oxygen being set free:

$$H_2O + Cl_2 = 2\ HCl + O.$$

Bottles in which chlorine-water is kept should, therefore, be protected from the light, and this can be most conveniently done by pasting black paper round them.

The bleaching and disinfecting power of chlorine is now easily explained by its strong affinity for hydrogen. All animal and vegetable substances contain hydrogen, which is taken from them by chlorine. But if a single chemical pillar falls, the whole chemical structure tumbles with it. By the abstraction of the hydrogen, the colouring matter becomes colourless, the odorous principles scentless, the morbific matter innoxious, insoluble substances are very frequently rendered soluble, &c.

Experiment 8.—Provide two bottles of chlorine, and place in one of them some dry calcium chloride. This salt eagerly absorbs water, and will thoroughly dry any gas in contact with it. In a few hours introduce a piece of blue litmus-paper into the dry chlorine; no change in the litmus will be apparent, or, at most, it will be but slightly reddened. Perform the same operation in the other bottle, and the litmus-paper will be rapidly bleached; proving that the presence of water is necessary to enable chlorine to exert its power in this respect.

Experiment 9.—Put into chlorine water some gold-leaf; it will soon disappear, as the element chlorine combines with the element gold. The combination is called *auric chloride;* it is soluble in water. Chlorine has a very *great tendency to combine with the metals.* These combinations comport themselves as salts; they are called *metallic chlorides*, and most of them are soluble in water.

Experiment 10.—Pour into a vessel filled with chlorine gas a little metallic antimony, in fine powder; it will fall in a red-hot state to the bottom, as though it were a shower of fire. The red heat is caused by the violent combination of the chlorine with the antimony. The white smoke which fills the flask is the new combination formed, viz., *antimonic chloride.* If a fine brass wire, on which a piece of tinsel has been fastened, be introduced into chlorine gas, both will burn with vivid combustion, and with the emission of fumes. Here combustion is another name for combination with chlorine. Brass consists of zinc and copper; accordingly, chlorides of zinc and copper are formed. Both dissolve in water, and the copper chloride imparts to the solution a green tinge.

Experiment 11.—If a piece of sodium of the size of a pea is thrown into a cup containing chlorine water, it will move rapidly round, just as in common water, with a hissing noise, and finally disappear; but if a sufficient quantity of the chlorine was present, the liquid will not afterwards give a basic reaction, as in Experiments 1 and 2, *Hydrogen;* neither will it have an alkaline, but a saline taste. If allowed to evaporate gradually over a warm stove, small cubic crystals remain behind, the constituents of which are chlorine and sodium. Thus, from these two elements a *salt* has been formed, familiarly known as *common salt*, NaCl.

Chlorine, like oxygen and sulphur, often unites in several proportions with a substance. Thus, there are two different chlorides of mercury.

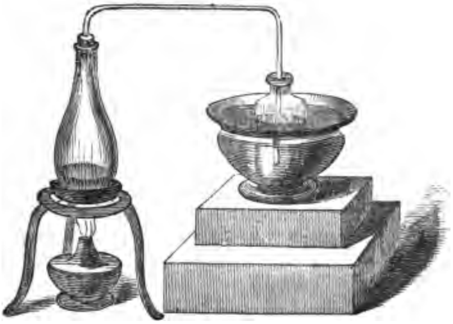
HYDROCHLORIC ACID, HCl.

Experiment 1.—Put into a porcelain capsule a few grains of common salt, and pour a little sulphuric acid upon it; there escapes, with effervescence, a gas, which has an acid taste, fumes in the air, and reddens moistened blue test-paper; this gas is *muriatic acid*, or *hydrochloric acid.* If you pour some ammonia upon a shaving, and wave the latter to and fro over the capsule, a thick white smoke is formed; and the odour of both the hydrochloric acid, and also the pungent fumes of the ammonia, vanish. The two gases combine and form solid *ammonium chloride:*

$$NH_3 + HCl = NH_4Cl.$$

Experiment 2.—Mix carefully in a flask a quarter of an ounce of water with three-quarters of an ounce of sulphuric acid, and after the mixture has become cold, add to it half an ounce of common salt. Adapt to the neck of the flask a cork provided with a glass tube, the long limb of which passes into a phial, containing one ounce of water. If you heat the flask in a sand-bath, the hydrochloric acid escapes, but more quietly than in the former experiment, because the sulphuric acid has been somewhat diluted. The tube must only just dip into the water; for should it reach to the bottom of the phial, the whole liquid might suddenly flow back into the flask, if the heat should chance to slacken, as it might, for instance, from the flickering of the lamp by an accidental current of air. The hydrochloric acid is so eagerly absorbed by the water, that, when the evolution of the gas diminishes, a vacuum is formed in the tube and flask; the pressure of the exterior air then forces the water up into the flask (page 54). When a gaseous body condenses into a liquid, it no longer requires the latent heat by which it became gas or vapour, and therefore this heat is set free. From this it follows that the water in which the muriatic acid condenses or dissolves must soon become warm. But warm water takes up much less gas than cold; accordingly, in order to obtain a concentrated solution of hydrochloric acid gas, we must place the phial in a basin of cold water. When the liquid in the receiver has sufficiently increased, one of the blocks must be withdrawn from beneath, so as to keep the end of the tube near the surface of the liquid. The solution thus obtained has an intensely acid taste and reaction; it is called

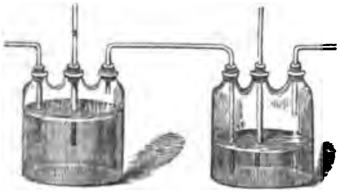
Fig. 43.



hydrochloric acid, but is often known by the name of *muriatic acid.* One measure of water absorbs more than four hundred measures of hydrochloric acid gas; the strong hydrochloric acid thus obtained fumes in the air, because a part of the gas escapes. If you heat it to boiling, then half of it escapes, and an acid only half as strong remains behind; but this is always somewhat *heavier* than water.

The hydrochloric acid of commerce is commonly yellow, and contaminated with sulphurous acid, sulphuric acid, chlorine, iron, and arsenic. It is manufactured from common salt and sulphuric acid; but, instead of glass vessels, large iron cylinders are employed, capable of containing some hundredweights of common salt. The gas is conducted into several bottles or jars filled with water, and connected with each other. When the water in the first vessel becomes saturated with the hydrochloric acid gas, the gas passes over into the second, then into the third vessel, and so on, saturating each successively. This is a very convenient method of conducting gases through liquids. Such vessels, which are commonly provided with two or three necks, are called *Woulfe's bottles.* The upright tube in the middle neck serves as a safety tube, that is, it prevents the liquid from being forced back; if a vacuum is formed in one of the bottles, the air enters through this tube.

Fig. 44.



Experiment 3.—If instead of dipping into water, the delivery-tube is passed to the bottom of a dry bottle, the heavy gas will displace the air, and fill the bottle. Its solubility may then be shown by inverting the bottle quickly in water. The water will rush up and fill the bottle.

Common salt consists of chlorine and sodium. With sulphuric acid the following change takes place:

$$\overset{\text{Sodium Chloride.}}{NaCl} + \overset{\text{Sulphuric Acid.}}{H_2(SO_4)} = \overset{\text{Sodium Sulphate.}}{Na_2(SO_4)} + \overset{\text{Hydrochloric Acid.}}{2HCl.}$$

The constituents of hydrochloric acid gas are equal atoms of chlorine and hydrogen, and it is represented by the formula HCl.

If you fill a glass vessel half with chlorine and half with hydrogen, and put it in a dark place, no union ensues; but it takes place instantaneously when a light is applied or when the vessel is exposed to the direct rays of the sun. The union is accompanied by a violent detonation, which often breaks the vessel, so that it is not advisable to perform this experiment except on a very small scale. In the diffused light of day the combination takes place slowly.

Experiment 4.—Put some iron nails into a phial, and pour upon them some hydrochloric acid; brisk effervescence will ensue. When this has continued some minutes, hold a burning taper over the mouth of the phial; the gas which escapes takes fire; it is hydrogen. The acid is decomposed, and its second constituent, chlorine, combines with the iron. The iron disappears, and is dissolved; that is, it combines with the chlorine, forming a soluble compound. When the effervescence has ceased, heat the phial by placing it in hot water, and afterwards pour its contents upon a filter of white blotting-paper. Put the liquid which passes through (*the filtrate*) in a cool place; a salt is deposited from it in greenish crystals, called *ferrous chloride*, $FeCl_2$.

Many other metals are also dissolved, like iron, by muriatic acid, and converted into chlorides.

Experiment 5.—Pour some muriatic acid upon iron-rust that has been put into a test-tube; it dissolves, but without evolution of gas. In this case, the hydrogen of the hydrochloric acid meets with a body with which it can combine, namely, the oxygen of the oxide or rust of iron; and it does combine with it, forming water. The yellowish-brown solution, which it is difficult to crystallise, yields, upon evaporation, a brown mass called *ferric chloride*, Fe_2Cl_6. This salt contains one half more chlorine than the former. Hydrochloric acid is very often used for dissolving metallic oxides.

Experiment 6.—Dissolve some crystals of the protochloride of iron, obtained according to Experiment 3, in a little water, and then add some chlorine water; the greenish colour is converted into a yellow colour, and the solution yields, on evaporation, brown *ferric chloride*. The chlorine converts the *ferrous* chloride into *ferric* chloride.

Experiment 7.—Dissolve some carbonate of soda (sodium carbonate) in water; the solution turns red test-paper blue;

it has a basic reaction. Drop carefully into the solution some hydrochloric acid, until neither the red nor the blue paper is affected by it. If you put the liquid in a warm place a salt will be deposited in small cubes; you readily perceive, both by the shape of the crystals and by the taste, that it is common salt. Carbonic anhydride escapes with effervescence.

Experiment 8.—To a little water in a test-tube add a drop of hydrochloric acid or a grain of a soluble chloride, such as common salt, and then a few drops of a solution of silver nitrate (lunar caustic); a white cloudiness is formed which does not happen in pure water. This cloudiness or *precipitate* is due to the formation of silver chloride, which is insoluble in water, but it may be dissolved by adding some *solution of ammonia.* Nitrate of silver is a most accurate test for hydrochloric acid or for chlorides:

$$AgNO_3 + NaCl = AgCl + NaNO_3.$$

BROMINE.

Symbol, Br = 80. Formula, Br_2.

Bromine is a deep brownish-red, heavy, and very volatile liquid. Its name is derived from the Greek word βρῶμος, signifying a disagreeable odour. Bromine, at common temperatures, emits yellowish-red fumes, which have a penetrating and offensive odour, resembling that of chlorine. It produces a yellow colour with starch, is sparingly soluble in water, but very soluble in ether. Its specific gravity is 3.

Bromine occurs in many mineral waters, some of which, as the Kreutznach spring, are comparatively rich in it, and form the source from whence it is obtained on the Continent. But in this country the chief supply is derived from sea-water. A large quantity of common salt is obtained by evaporating sea-water, and it is from the residual liquid, which is technically called *bittern*, that bromine is obtained. The bittern is mixed in a retort with manganese peroxide, and hydrochloric acid. On the application of heat the chlorine liberated from the acid decomposes the magnesium bromide contained in the bittern, and bromine, mixed with water, distils over:

$$MgBr_2 + Cl_2 = MgCl_2 + Br_2.$$

The bromine is separated from the water by shaking it with ether. The ethereal solution, which has a brown colour, floats on the water and can be drawn off. The bromine is then extracted by a somewhat complex process.

Experiment 1.—The processes for the preparation of bromine may be imitated by the student by dissolving a few grains of potassium bromide in a little water contained in a long test-tube; this may represent the bittern. A little chlorine water will liberate the bromine, and the solution will become brown. If a drachm or two of ether is then added and the mixture violently shaken for a minute, the ether will remove all the bromine, and on standing a little while will float as a dark layer on the surface of the water. Pour off the dark-coloured ether into a porcelain basin, and add a few drops of potassium hydrate (caustic potash); it will become colourless. Heat it gently until all the liquid is driven off and a white substance will remain behind, which on cooling may be mixed with a few grains of manganese peroxide, and two or three drops of sulphuric acid, and again heated. Free bromine will then be evolved as brown vapours, very readily seen against the white sides of the basin.

Experiment 2.—Pour a few drops of bromine into a small phial with some distilled water, and shake the mixture well. The bromine will partly dissolve and form a brown solution. This may be preserved and labelled "Bromine solution."

Experiment 3.—Add a few drops of the above solution to a little very weak mucilage of starch, made by mixing a few grains of starch with a *little* cold, and then adding a considerable quantity of boiling, water with continual stirring. A pale yellow colour will be produced, owing to the formation of bromide of starch.

Experiment 4.—To a drop of bromine contained in a wine-glass add a *minute* fragment of phosphorus, which must be cut off under water, as it is very inflammable, and carefully and quickly dried in a piece of blotting-paper. The best plan is to drop the phosphorus from the point of a knife on to the bromine. The phosphorus will be immediately inflamed, and a smart explosion will occur.

IODINE.

Symbol, I = 127. Formula, I_2.

Iodine is a black solid of metallic lustre; it smells somewhat like chlorine, has a pungent taste, and stains the skin brown.

It occurs in small quantity in sea-water, from which it is separated and assimilated by the various sea-weeds, which thus accumulate a considerable quantity; and it is from the ashes of certain kinds of sea-weeds that iodine is always obtained, for sea-water contains much too small a quantity to render its extraction profitable. The sea-weed is burnt and the ash, or *kelp*, as it is called, contains the iodine in the forms of magnesium iodide, MgI_2. and sodium iodide, NaI. The kelp, after some preliminary treatment, is heated with black oxide of manganese and sulphuric acid, and yields iodine just as the chlorides or bromides yield chlorine and bromine respectively when treated in the same way.

Experiment 1.—To a few grains of potassium iodide, or some other iodide, in a test-tube, add an equal weight of black oxide of manganese, and a drop or two of sulphuric acid, and heat gently. Beautiful violet vapours of iodine will be produced, and will condense on the sides of the test-tube as black, shining spangles. The following change occurs:

$$2KI + MnO_2 + 2H_2SO_4 = K_2SO_4 + MnSO_4 + 2H_2O + I_2.$$

Experiment 2.—Dissolve a grain or two of potassium iodide in a little distilled water in a test-tube, and add a little solution of silver nitrate. A light yellow precipitate (silver iodide) will be thrown down, which will not be redissolved on the addition of solution of ammonia, thus clearly distinguishing it from the similar white precipitate obtained with chlorides.

Experiment 3.—Put 24 grains of iodine into a flask, and pour over them half an ounce of strong alcohol; if the iodine is pure it will entirely *dissolve*. This dark brown solution is called *tincture of iodine*. Water dissolves only a trace of iodine, but yet is rendered yellow by it.

Experiment 4.—Put a little iodine upon a knife, and hold it over the flame of a lamp; the iodine melts, and is afterwards converted into a *violet-coloured gas—vapour of iodine.*

As the iodine fumes are nearly nine times heavier than common air, they sink in it. Iodine owes its name to the colour of its vapour, the Greek word ἰώδης meaning *violet-coloured.* The vapour appears more beautiful when the iodine is heated in a small flask. After cooling, the walls of the flask become lined with small brilliant crystals of solid iodine, affording an example that regular crystals may be formed when bodies pass from the aeriform into the solid state.

Experiment 5.—Boil one grain of starch in a test-tube with one drachm of water, and add to the thin paste thus obtained a few drops of tincture of iodine; *the iodine combines with the starch; the combination is of a deep blue colour.* The blue colour disappears on careful heating, but returns again on cooling. If one drop of the starch paste is mixed with one quart of water, even at this extreme dilution, the iodine tincture will impart to it a violet tinge. Consequently it is an exceedingly delicate test for detecting starch, and starch, on the other hand, for detecting iodine. If a little tincture of iodine is dropped upon flour, bread, potatoes, &c., the presence of starch in these substances will at once be indicated.

COMPOUNDS OF BROMINE AND IODINE WITH HYDROGEN.

Hydrobromic Acid, HBr, *and Hydriodic Acid*, HI.—Both of these acids closely resemble hydrochloric acid. They are colourless gases, which fume strongly in air, and are readily soluble in water, forming acid solutions. With bases they form iodides and bromides:

$$KHO + HBr = H_2O + KBr.$$

$$KHO + HI = H_2O + KI.$$

Both bromine and iodine are faithful companions of chlorine; wherever common salt occurs, whether in the earth, the sea, or mineral springs, small quantities of them are present, not in a free state, however, but combined with metals. The different sea-weeds attract these combinations from the sea-water, and from these sea-weeds iodine and bromine are extracted. Both have poisonous properties.

FLUORINE.

Symbol, F = 19. Formula unknown.

Fluorine is unknown in its isolated state. When liberated from its compounds it acts energetically on all substances of which vessels are ordinarily made. The mineral known as fluor-spar consists of fluorine and calcium.

Hydrofluoric Acid, HF.

Experiment 1.—Fashion a small basin out of a piece of thin sheet lead, place in it about a quarter of an ounce of powdered fluor-spar, and half an ounce of strong sulphuric acid, and apply a gentle heat. Fumes of hydrofluoric acid will arise, which are extremely irritating to the eyes, and which resemble those of hydrochloric acid. Fluor-spar (calcium fluoride) is decomposed by sulphuric acid just as common salt is:

$$CaF_2 + H_2SO_4 = CaSO_4 + 2HF.$$

Hydrofluoric acid is intensely corrosive, and it attacks and destroys glass and earthenware and most other materials. But it does not act upon lead; therefore vessels made of that material must be used to prepare it in. The most interesting property of hydrofluoric acid is its power of corroding glass, and it is often used in the arts to etch designs upon glass, which is done in the following manner.

Experiment 2. — Thoroughly warm a small square of window glass, and then rub over one side of it with a piece of wax, so as to make a smooth coating. When the glass has become cold and the wax hard, scratch any design with a pointed instrument on the coated side of the plate, taking care to cut quite through the wax. The glass plate thus prepared is then laid with the waxed side downwards, so as to cover the leaden basin in which the hydrofluoric acid is being prepared as directed in Experiment 1.

If, after being exposed to the acid fumes for a few minutes, the plate is removed and cleaned from the wax with a rag and a little turpentine, it will be found corroded wherever the marks have been made, and the design will remain engraved on the glass.

CHAPTER II.

NON-METALLIC DIADS.

(Oxygen, Sulphur, Selenium, Tellurium.)

OXYGEN.

Symbol, O = 16. Formula, O_2.

Oxygen is the most abundant of the elements, for it is probable that not less than two-thirds of all that portion of our globe which is known to us consists of it. About one two-billionth part of this oxygen exists in the atmosphere in a free, or uncombined state ; the remainder occurs in combination with various other elements. It constitutes eight-ninths of the weight of water, and about one-half of the weight of the chief rocks, and it is an important constituent of all animals and vegetables.

In the pure state it is a gas of specific gravity 16, that is, about one-tenth heavier than common air. It is colourless, tasteless, and inodorous, and is very slightly soluble in water. Its compounds are called *oxides*, and oxides of all the elements, except fluorine, are known. The gas is magnetic; for a balloon filled with it is attracted by the poles of a powerful magnet.

Oxygen cannot easily be prepared from air, and in the great majority of its most abundant compounds it is held in combination by a force too strong to be easily overcome; but a certain number of oxygen-compounds are known, which are so unstable, that either by heat, or by some other agency, the element may be displaced and collected in a pure state.

PREPARATION.

Oxygen was first prepared by Dr. Priestley, in 1774, from mercuric oxide, by the process already described (page 61).

It is also obtained when water is decomposed by electricity (page 12).

Experiment 1.—Into the same apparatus that was used for heating mercuric oxide, introduce 100 grains of *potassium chlorate*, and heat it carefully with the spirit-lamp. The salt will soon melt, and afterwards boil. As soon as the boiling commences the flame must be lowered, to prevent the mass from frothing over. When the liquid thickens, if some of the substance should be found adherent to the colder parts of the tube, approach it with the flame of the lamp until it is again melted down. The gas may be collected over the pneumatic trough, just like hydrogen ; but when it ceases to come off, the delivery tube must *immediately* be removed from the water.

Instead of using pure potassium chlorate, it is much better to mix the salt with about half its weight of *carefully-dried* manganese peroxide. The oxygen is then given off at at a lower temperature and more easily, but the manganese peroxide remains unaltered after the experiment.

The following equation shows the change which takes place in this important reaction :

$$\underset{\text{Potassium Chlorate.}}{2KClO_3} = \underset{\text{Potassium Chloride.}}{2KCl} + 3O_2.$$

Potassium chlorate contains for every one hundred grains nearly forty grains of oxygen chemically combined ; by the application of heat, these become free and escape. Mercuric oxide contains nearly eight per cent. of oxygen ; therefore the former will yield five times more oxygen than the latter. If phials of twelve ounces' capacity are selected for receiving the gas, we shall be able to fill five of them, and shall have in each about eight grains, or nearly twenty cubic inches, of oxygen.

Potassium chlorate may, under some circumstances, as when strongly rubbed, or treated with sulphuric acid, occasion *very dangerous explosions ;* but no danger is to be apprehended from the application of it in the manner above directed.

Experiment 2.—Add warm water to the salt remaining in the test-tube after the expulsion of the oxygen, and place the tube in a warm place until the salt is dissolved ; evaporate the solution gradually, when small cubic crystals of potassium chloride will be deposited. The potassium chlorate crystallises in thin tables or plates, the heated mass in cubes; this difference in the form of the crystals indicates that, by the heating of the former, an entirely new salt is formed. It is indeed one which no longer contains oxygen.

Oxygen may also be prepared by the following processes, which, however, are less convenient on the small scale than the foregoing:

1. By heating manganese peroxide to redness in an iron tube closed at one end:

$$3MnO_2 = Mn_3O_4 + O_2.$$

2. By passing the vapour of sulphuric acid through a red-hot tube:

$$\underset{\text{Sulphuric Acid.}}{2\,H_2SO_4} = 2\,H_2O + \underset{\text{Sulphurous Anhydride.}}{2\,SO_2} + O_2.$$

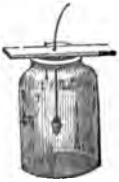
The sulphurous anhydride may be absorbed by passing the gas through lime. This, on the large scale, is the cheapest process for preparing oxygen.

PROPERTIES.

Experiment 3.—Introduce a glowing shaving into a bottle of oxygen; it will kindle and burn for some time with great brilliancy and with a very dazzling flame, and then be extinguished. The same takes place when a piece of lighted tinder is fastened to a wire and suspended in the oxygen; the tinder burns with a lively flame, while, as is well known, it merely smoulders away in the open air. Oxygen, at a high temperature, combines eagerly with the component parts of wood and tinder, and heat and light are developed. When the combination is ended, and the oxygen is consumed, the combustion ceases. The product of the combustion, that is, the combination of the wood with the oxygen, is also aeriform; but burning substances are extinguished in the newly-formed gas. If the bottle be rapidly whirled round, the gas formed by the combustion will escape, and atmospheric air will supply its place. Air contains free oxygen; and a kindled shaving will burn in it for some time, but far slower and less briskly than in pure oxygen; because common air contains only one-fifth part of oxygen. Accordingly, combustion proceeds much more rapidly and violently in oxygen than in atmospheric air.

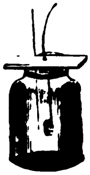
Experiment 4.—Fasten a piece of charcoal to a wire, and kindle it in the flame of a lamp, and then introduce it into a bottle of oxygen; it will burn very vividly, and, if the charcoal had much bark on it, with beautiful sparks. If a piece of moistened blue litmus-paper be introduced into the bottle, after the combustion, it will be reddened; consequently an acid oxide has been formed from the charcoal and the oxygen; it is called *carbonic anhydride.* Pour into the bottle a little *lime-water* (which is prepared by shaking slaked lime with water for a short time and filtering). The clear lime-water becomes milky from the formation of the insoluble calcium carbonate. This is a good test for carbonic anhydride.

Fig. 45.



Experiment 5.—If some pieces of sulphur are fastened to a longer wire, kindled and suspended in a second bottle, they will burn with a beautiful blue flame. The gas formed from this union of sulphur and oxygen has a very irritating odour; when dissolved in water, it likewise turns litmus-paper red, and consequently it is of an acid nature. It is called *sulphurous anhydride.*

Fig. 46.



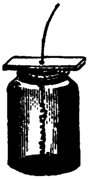
Experiment 6.—Take a small piece of phosphorus, which, on account of its inflammability, must be cut off under water from the stick, and place it, after it has been *well dried* between blotting-paper, in a scooped-out piece of chalk. Fasten the latter to a wire, and introduce it into a third bottle of oxygen. Affix the wire to a cross piece of wood, so that the chalk may hang a little below the centre of the bottle. If the phosphorus be now touched with a hot wire, it will kindle and burn with a dazzling brilliancy, filling the bottle with a thick white smoke. This smoke consists of a chemical compound of oxygen and phosphorus; it reddens the blue test-paper, and consequently is also an *acid oxide;* it is called *phosphoric anhydride.* If the bottle be allowed to stand for a time, the smoke will sink to the bottom and dissolve in the water which remains there, which thus acquires an acid taste.

The two last experiments, and many others, can be more conveniently performed in *deflagrating spoons*, which are merely little ladles of iron or brass fastened at right angles on the end of iron wires, which pass through a cork and a piece of tin plate. The tin plate rests on the mouth of the bottle.

Experiment 7.—Heat a small piece of sodium in a deflagrating spoon till it begins to burn, and then plunge it in a bottle of oxygen. It will burn with great brilliancy, and with a yellow flame. A white oxide, called *sodium oxide*, Na_2O, is formed, which, when dissolved in water, turns red litmus-paper to blue. Oxides of this kind are called *basic oxides* (page 98). Potassium yields a similar compound.

Experiment 8.—A piece of fine iron wire is so wound round a slate- or common lead-pencil, that, on the withdrawal of the latter, the wire may have a spiral form. Fasten the upper part of this wire, as in Experiment 6, to a cross-piece of wood, and place on the lower end of it a small portion of tinder. When this is kindled, introduce the wire into the oxygen; the burning tinder heats the iron to redness, which then burns brilliantly, throwing out sparks. The iron, when red-hot, combines with the oxygen. The burnt or oxidized iron (iron scales) melts, and falls to the bottom in black globules, which are so hot that they are apt to melt into the glass, though it be partly filled with water. This black oxide of iron, Fe_3O_4, is the same that is formed when steam is passed over red-hot iron in a tube. The inside of the bottle becomes covered during the experiment with a brick-red oxide, Fe_2O_3, which is called *ferric oxide*. Common iron rust is ferric oxide. Ferric oxide is a basic oxide, although, being insoluble in water, it cannot be tested with litmus-paper. The black oxide is a *neutral* oxide. It is neither acid, nor basic.

Fig. 47.



These experiments show that there are three kinds of oxides:

1. Acid oxides, or anhydrides, which by combination with water yield acids.

2. Basic oxides, which by combination with water yield bases.

3. Neutral oxides.

The connection that these oxides have with the acids and bases has already been described (page 97).

Combustion.—The process which we call combustion, or burning, is nothing but rapid combination with oxygen, with evolution of light and intense heat. It is in fact only a particular case of oxidation. In all ordinary cases oxidation is attended with the production of heat, which we have before seen (pp. 13, 27) is a frequent result of the exertion and consequent loss of chemical force. But the oxidation of a certain weight of substance may take place rapidly or slowly, and the *intensity* of the heat will, of course, vary in like proportion, although the total *quantity* remains the same. The term combustion is generally limited to those cases of oxidation where light as well as intense heat is evolved. But the term may also be applied to similar combinations between other elements, as, for instance, between chlorine and hydrogen or metals.

Experiment 9.—Invert a tumbler over a small piece of lighted candle standing on a plate. In a short time the candle will go out, having removed a portion of the oxygen from the air of the tumbler. The inside of the tumbler will be seen to be covered with *moisture*, and if it is rapidly removed, and a little lime-water shaken in it, the latter will become milky, showing the formation of carbonic anhydride.

All ordinary kinds of fuel—coal, wood, gas, oil, &c.—contain carbon and hydrogen. This experiment shows the important fact that, *when they burn, the carbon becomes carbonic anhydride, and the hydrogen water.*

The phenomena of slow oxidation are very interesting and important.

Experiment 10.—Throw some iron filings, which have been moistened with water, into a bottle, so that they may adhere to its bottom. Invert the bottle with its mouth in a basin of water, and leave it in a warm place for a few days. The water will gradually ascend in the bottle, for the air becomes diminished in bulk to the extent of one-fifth; in fact, the whole of the oxygen is slowly absorbed by the iron and combines with it, forming *ferric oxide*, Fe_2O_3. The gas remaining in the bottle is *nitrogen*, as is shown by plunging a lighted taper into the gas; the taper will be extinguished (see Nitrogen). The *rusting* of iron is here seen to be its slow oxidation.

Experiment 11.—Bruise a little hay with water, stuff it in the bottom of another small phial, and proceed as in the former experiment. The same effect will be produced as with the iron filings; that is to say, the oxygen will be absorbed by the hay, and nitrogen will remain, mixed in this case with a little carbonic anhydride from the oxidation of the hay. The hay undergoes the process of *decay*, which is thus shown to be a slow oxidation. In both of these cases, heat is evolved, though it possesses but slight intensity. The heat of a *hot-bed* is due to the slow oxidation of the manure.

Respiration.—Experiment 12.—The heat of the animal body is maintained by a process of slow oxidation, analogous to that of the hay. Blow through a glass tube into a tumbler half full of lime-water. In a few minutes it will become turbid, showing that *carbonic anhydride* is contained in the expired air. Breathe into a cold and dry tumbler. It will become covered with *moisture*. Animals consume the oxygen of the air, and return carbonic anhydride and water to it.

Substances are more easily burnt when they are in a finely divided state than if they are solid; shavings are easier to burn than billets of wood, and iron is easily combustible when in the form of filings.

Experiment 13.—Hold a small bar of iron in the flame of a spirit lamp; it will not burn; but file it and let the filings fall into the flame; they will be immediately ignited. Iron filings when sprinkled in a flame burn brilliantly, and have a very pretty appearance.

Deflagration.—The term *deflagration* is applied to those cases of combustion where the body is burnt, not by the agency of *free* oxygen as that of the air, but by the aid of *combined* oxygen as it exists in substances which are easily decomposed, such as potassium chlorate, $KClO_3$, and potassium nitrate, KNO_3. Deflagration may be exemplified as follows:

Experiment 14.—Heat 20 or 30 grains of nitre in a test-tube with a spirit lamp, and when it is melted throw in some powdered charcoal or a fragment of sulphur; it will burn violently, being oxidized at the expense of the nitre.

OZONE, O_3.

When certain kinds of electrical discharge occur in an atmosphere of pure oxygen, or even of air, a contraction of volume takes place, and some of the oxygen is converted into the curious gas called ozone (Gr., ὄζω, to smell), on account of its peculiar odour. This odour can be perceived in the neighbourhood of an electrical machine during its action. The specific gravity of ozone is found to be 24 ($H = 1$), that of oxygen being 16; so that ozone is half as heavy again as oxygen. This fact is explained by the more than probable theory that every molecule of ozone contains three atoms of oxygen instead of two, as in ordinary oxygen gas.

Ozone has never been obtained free from oxygen. It is an amazingly powerful oxidizing, bleaching, and disinfecting agent, a property which is due to the ease with which it gives out its third atom of oxygen. At a temperature of about 250° C. (482° F.), it is reconverted into common oxygen, $2O_3 = 3O_2$. It can be formed by several methods besides the one above mentioned.

Experiment 1.—Take a stick of phosphorus, about an inch long, and scrape it with a knife, under water, so as to expose a new surface; then lay it in the bottom of a large wide-mouth bottle, pour in enough water to half cover it, and having loosely stoppered the bottle, set it aside in a warm place for about half an hour. White fumes of phosphorous acid will arise from the phosphorus, but will soon be absorbed by the water, and, simultaneously, ozone will be formed. The phosphorus combines gradually with part of the oxygen of the air in the bottle, while some of the remainder is converted into ozone, $3O_2$, becoming $2O_3$. Its presence may be shown by the use of what is called the "ozone test," which is made by dissolving a few grains of potassium iodide in a table-spoonful of starch jelly, such as is prepared for use in the laundry, and smearing the latter on strips of paper. Introduce one of these test-papers in the bottle of ozone, and it will immediately become dark blue or violet, the colour being more or less intense according to the quantity of ozone. The iodine is liberated from the potassium iodide by the ozone, and it combines with the starch to form iodide of starch (Experiment 5, page 119).

Experiment 2.—Lay a piece of the test-paper in the bottom of a tumbler and gradually invert a bottle of ozone, prepared as in Experiment 1, as though in the act of pouring from the one vessel to the other. The ozone will stream out and occupy the tumbler, and come in contact with the test-paper, which will be immediately affected. This experiment shows that ozone is heavier than air.

Experiment 3.—Suspend a bright silver coin in a bottle of ozone. In a few minutes it will be covered with a grey deposit of silver oxide, thus showing the remarkable oxidizing power of ozone.

Experiment 4.—Place a few fragments of sulphide of iron in a glass or jar, and pour upon them a little water and a few drops of sulphuric acid; sulphuretted hydrogen, a gas of an extremely disagreeable odour, will be evolved. Moisten a strip of paper with a solution of lead acetate, and bring it in contact with the escaping sulphuretted hydrogen, and the paper will become dark brown or black, from the formation of *lead sulphide*, PbS. If the paper be then plunged into a bottle containing ozone it will again become colourless. The ozone oxidizes lead *sulphide*, which is black, into lead *sulphate*, which is white:

$$PbS + 4O_3 = PbSO_4 + 4O_2.$$

Experiment 5.—Moisten a glass rod with a strong solution of ammonia, and introduce it into a bottle of ozone. White fumes will be abundantly formed, which consist of *ammonium nitrite*, NH_4,NO_2.

Experiment 6.—Moistened litmus-paper is immediately *bleached* when introduced into a bottle of ozonized air. The bleaching and disinfecting of bodies by ozone are owing to their oxidation. Ozone is very commonly found in the atmosphere, formed perhaps by the passage of lightning through it. It cannot be doubted that this atmospheric ozone has important functions to fulfil in the economy of nature.

Experiment 7.—Strips of paper moistened with ozone test may be exposed to the air for a few hours. They will frequently be found to have turned blue, especially in country places.

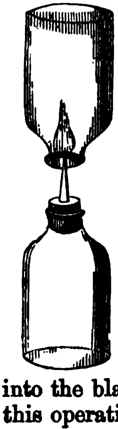
WATER, H_2O.

Water does not exist in nature in a perfectly pure condition, even fresh water contains small quantities of solid impurities, of which the carbonates and sulphates of calcium and magnesium are among the most important and general. Natural water also holds in solution small quantities of oxygen, nitrogen, and carbonic acid; but besides these general impurities, there are others which occur only in particular instances, such as iron, sulphuretted hydrogen, iodine, &c., which are present in certain springs, and which give to them a medicinal value, as is the case with the waters of Harrogate, of Cheltenham, of Tunbridge Wells, &c. Rain water, collected in a vessel placed in a large open space, such as a field, is the purest kind of *natural* water that can be obtained, and may be used by the student, for most purposes, in lieu of *distilled* water, in cases where the latter cannot easily be obtained.

The composition of water was first determined by Cavendish; he proved that it is the sole product of the combustion of hydrogen.

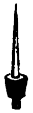
Experiment 1.—That water is really formed by the burning of hydrogen can easily be shown by inverting a clear glass bottle over the hydrogen flame; the glass soon becomes clouded over, because the water, which in consequence of the heat is generated in the form of steam, condenses in small globules on the cold sides of the bottle.

Fig. 48.



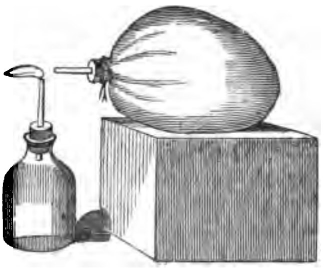
Experiment 2.—The extraordinary degree of heat developed by the chemical union of oxygen and hydrogen may be shown by the following experiments. Insert into the opening of a large pig's bladder, which has been softened by soaking in water, the broken-off neck of a flask, and bind it firmly round with a string. Then select two perforated corks, fitting this neck. One cork is connected with a bent glass tube, conducting the oxygen from the apparatus in which it is evolved into the bladder, which soon becomes filled with it. When this operation is finished, replace the first cork by the second, having a glass tube adapted to it only a few inches long, and drawn out to a point at its outer end, and provided with a wax stopple pressed upon the opening. A glass tube may be formed into a jet by heating it in the flame of a spirit-lamp, constantly turning it round at the same time, till it becomes so soft at the desired place as to be easily drawn out. Break it at the slender part, and hold it in the flame for some moments, until the sharp edges are rounded off by incipient melting. It would be more convenient, though somewhat more expensive, to substitute for the above contrivance a small brass stop-cock, provided with a jet.

Fig. 49.



The bladder thus arranged and filled with oxygen is now placed upon a block, at such a height that the point of the glass tube shall be on a level with the hydrogen flame, produced as explained in a former experiment. Press upon the bladder with the hand, and the oxygen will escape, blowing into the hydrogen flame, which then takes a horizontal direction. This flame has but little brilliancy, less than the hydrogen flame alone, notwithstanding which it affords the greatest heat yet known, except the heat produced by electricity. Hold in it a platinum wire, a metal which cannot be melted in the hottest furnace, and it will melt like wax; hold in it a piece of lime scraped to a fine point, and it will emit light of the most dazzling splendour. This is called the *lime light.* A watch-spring or a fine iron wire burns in it, throwing out sparks as in oxygen.

Fig. 50.



Exact experiments have shown that two measures of hydrogen unite with one measure of oxygen, the same proportion as that in which the gases are obtained when water is decomposed by electricity (page 12). The result of the combination is steam, which condenses into water. But two measures of hydrogen and one of oxygen do not yield three measures of steam; they afford two measures only:

$$\underset{\text{(4 Vols.)}}{2H_2} + \underset{\text{(2 Vols.)}}{O_2} = \underset{\text{(4 Vols.)}}{2H_2O}.$$

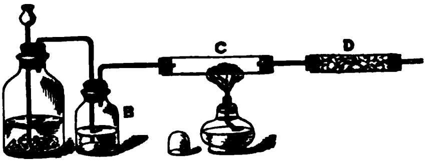
Thus the two gases condense one-third by chemical union.

If the hydrogen and oxygen were mixed together and then ignited, the whole mass would combine together at once, producing a most violent report, and bursting the vessel to pieces. The explosion is due to the concussion of the air, caused by the sudden and violent expansion of the steam by the heat evolved. No danger is to be apprehended from the apparatus described, as the explosive gas is formed at the point where the oxygen meets the hydrogen flame, and only in small quantities at once. The apparatus is an *oxy-hydrogen blowpipe* on a small scale.

Experiment 3.—Procure a soda-water bottle fitted with a good cork, and having filled it with water, invert it in the pneumatic trough, and fill it two-thirds full of hydrogen, and one-third with oxygen; cork it securely, and allow it to remain with its mouth immersed in the water for five minutes, so that the gases shall thoroughly intermix. If the bottle be then uncorked, and a light quickly applied to the mouth, the gases will combine with a loud explosion. Owing to the small quantity of the gaseous mixture employed there is no danger in this experiment; but to prevent a chance of accident it is advisable, before applying the light, to fold the bottle in a cloth, so as to secure the fragments of glass, should the bottle be broken by the explosion: such a result, however, is unlikely to occur.

The composition of water by volume is thus easily illustrated; but to show directly the relative *weight* of its constituents, the following rather complicated arrangement may be used:

Fig. 51.



A is a hydrogen apparatus, and it is connected by a bent tube with the bottle B, containing sulphuric acid. B is connected in turn to the two wide tubes C and D. C contains a small heap of black copper oxide, CuO, heated with a spirit lamp, and D is filled with fragments of calcium chloride.

Some granulated zinc having been placed in the bottle A, a mixture of sulphuric acid and water is poured down the long funnel called a thistle funnel. Hydrogen is evolved and passes by the bent tube through the sulphuric acid in the bottle B, and is thus dried, for sulphuric acid eagerly combines with and retains any watery vapour that may be carried over by the hydrogen. The dry gas, after bubbling through the sulphuric acid, passes through the wide tube C, and comes in contact with the copper oxide. *When the common air is expelled*, the tube C is heated. The hydrogen then combines with the oxygen of the copper oxide and forms water: $CuO + H_2 = Cu + H_2O$. The latter, as fast as it is formed, is driven forward and absorbed by the calcium chloride—a substance already mentioned as a powerful absorbent of water—in the tube D. Now if the tube D be weighed before and after the experiment, any additional weight it may have gained will of course be owing to the water formed during the experiment, and thus its amount may be readily determined; but if the tube C containing the oxide of copper be also weighed, any *decrease* in weight will represent the quantity of oxygen which has been required to form the water obtained. In this way it may be determined that for every *nine* grains of water produced, the oxide of copper loses *eight* grains of oxygen: the difference is of course due to the hydrogen, and represents the relative weight with which it combines with oxygen to form water. At the end of the experiment the tube C contains metallic copper, in fine powder.

Hydrogen Peroxide, H_2O_2.—Hydrogen may be made by indirect means to combine with an additional atom of oxygen than that required to form water. This compound is an oily liquid, easily decomposed, and of extreme oxidizing powers, rivalling those of ozone.

COMPOUNDS OF OXYGEN AND CHLORINE.

Oxides of Chlorine.		Corresponding Acids.	
Cl_2O	Hypochlorous Anhydride.	Hypochlorous Acid,	$HClO$.
Cl_2O_3	Chlorous Anhydride.	Chlorous Acid,	$HClO_2$.
ClO_2	Chlorine Peroxide.		
Cl_2O_5	*Chloric Anhydride?*	Chloric Acid	$HClO_3$.
Cl_2O_7	*Perchloric Anhydride?*	Perchloric Acid	$HClO_4$.

It is not necessary for the beginner to study all the above compounds. Many of them are highly dangerous substances, and all are difficult to prepare. The two anhydrides printed in italics have not yet been prepared, though the corresponding acids are well known. Chlorine peroxide is not an acid oxide.

Hypochlorous Anhydride, Cl_2O, is a colourless gas, difficult to prepare, and easily condensable to a red explosive liquid by cold. Combined with water it yields hypochlorous acid:

$$Cl_2O + H_2O = 2HClO.$$

This acid corresponds to the series of salts called *hypochlorites:*

$H(ClO)$, Hypochlorous Acid.
$Na(ClO)$, Sodium Hypochlorite.
$Ca''(ClO)_2$, Calcium Hypochlorite.

The substance commonly called chloride of lime, or bleaching powder, is intermediate in composition between the chloride and hypochlorite of calcium ($CaCl_2$ and $Ca(ClO)_2$). Its composition may be represented by the formula $Ca''Cl(ClO)$. It affords an example of the combination of an element with two distinct radicles. Bleaching powder is prepared by acting on slacked lime (calcium hydrate) with chlorine.

Experiment 1.—Throw a little slacked lime into a bottle of chlorine, stoppered loosely. In a short time the chlorine will disappear, the lime will absorb it, and be converted into bleaching powder:

$$\underset{\text{Calcium Hydrate.}}{Ca(HO)_2} + Cl_2 = \underset{\text{Bleaching Powder.}}{CaCl(ClO)} + H_2O.$$

The manufacture of bleaching powder on a large scale is carried on by passing chlorine through boxes filled with perforated trays, on which layers of slacked lime are spread. The process is complete when the lime ceases to absorb chlorine.

Experiment 2.—Pour a few drops of vinegar or diluted sulphuric acid on a little bleaching powder. Effervescence will be produced by the escape of gaseous hypochlorous acid, which has an odour somewhat resembling that of chlorine. A piece of moistened litmus-paper, if it be immersed in the jar, will be bleached. If sulphuric acid is added *in excess*, chlorine is set free:

$$CaOCl_2 + H_2SO_4 = CaSO_4 + H_2O + Cl_2.$$

This is a simple and effectual way of preparing chlorine for disinfecting purposes.

In the ordinary process of bleaching, the materials to be bleached are dipped into a solution of bleaching powder, and then passed through a dilute acid ("soured"), which liberates the hypochlorous acid. The acid is an even more powerful bleaching agent than chlorine.

The peculiar odour which bleaching powder apparently possesses is due to hypochlorous acid, which is slowly and continuously formed from it by the carbonic acid of the air, so easily is the compound decomposed.

Chlorine peroxide, ClO_2 or Cl_2O_4, is a dark yellow, easily decomposed gas, formed by the action of sulphuric acid on potassium chlorate. It may be prepared, and its explosive character safely observed, in the following way:

Experiment 3.—To two or three grains (not more) of potassium chlorate in a test-tube add a few drops of strong sulphuric acid, and gently warm. A peculiar crackling noise will be heard, while a coloured gas is slowly evolved. This gas is chlorine peroxide. If the tube be more strongly heated a loud explosion will often be produced from the sudden decomposition of the gas.

Chloric acid, $H(ClO_3)$, an oily liquid, easily decomposed, and of little utility. Its salts, especially the potassium salt, are of great importance. Chloric acid can only be prepared from its compounds, which are themselves readily obtained. By acting upon barium chlorate with sulphuric acid, chloric acid is formed, together with the insoluble substance, barium sulphate:

$$\underset{\text{Barium Chlorate.}}{Ba\,(ClO_3)_2} + \underset{\text{Sulphuric Acid.}}{H_2SO_4} = \underset{\text{Barium Sulphate.}}{BaSO_4} + \underset{\text{Chloric Acid.}}{2HClO_3}.$$

Potassium chlorate is formed by a process analogous to that used for the preparation of bleaching powder. By passing chlorine through a cold dilute solution of potassium hydrate, a mixture of potassium chloride and hypochlorite is obtained; but if a hot solution be used, chlorate instead of hypochlorite is formed:

$$6KHO + 3Cl_2 = 5KCl + KClO_3 + 3H_2O.$$

Experiment 4.—Pour a few drops of a strong and warm solution of potassium hydrate into a bottle of chlorine loosely stoppered. The chlorine will be absorbed, just as it was by calcium hydrate, while a white crystalline powder is abundantly formed which mainly consists of potassium chlorate, the chloride being very soluble. The chlorates are unstable bodies, and readily yield up their oxygen when heated; they therefore furnish convenient sources of that gas. It will be remembered that potassium chlorate is the salt commonly employed for the preparation of oxygen.

Experiment 5.—Powder a few grains of potassium chlorate; mix it gently with an equal quantity of powdered sugar, and touch the mixture with a rod dipped in strong sulphuric acid. A brilliant combustion will ensue, owing to the ClO_2 set free from the chlorate.

Experiment 6.—If *one* grain of potassium chlorate is triturated in a mortar with one grain of sulphur, a sharp detonation will follow, the sulphur being oxidized at the expense of the salt.

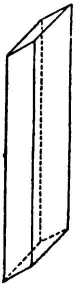
SULPHUR.

Symbol, $S = 32$. Formula of Vapour, S_2 and S_6.

The well-known element sulphur, which, on account of its easy combustibility, is employed in the manufacture of matches, &c., has neither taste nor smell. It is found in nature in considerable quantity, mostly in a state of combination.

Experiment 1.—Sulphur is *fusible*. Heat two ounces of sulphur in small pieces in a small stone-ware crucible or Florence flask, over a lamp at a gentle heat; it is converted, at a temperature a little above that of boiling water, into a thin, yellow fluid. If you pour some of it into cold water, you obtain again solid sulphur. If this, after being previously dried, is returned to the crucible, it will sink in the fluid mass, showing that solid is heavier than melted sulphur. Almost all other bodies behave in the same manner; ice, which floats on water, being an exception.

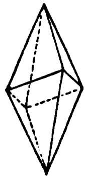
Fig. 52.



Experiment 2.—Sulphur may be crystallised. Let the crucible containing the melted sulphur stand till a crust has formed over the surface; break this quickly, and pour out the portion remaining fluid. Upon afterwards breaking the crucible, the cavity of the sulphur will be found lined with fine transparent crystals, in the form of lengthened pillars (Fig. 52), which are called oblique rhombic prisms.

In different parts of the world, particularly in volcanic countries, large beds of sulphur (native sulphur) are not unfrequently found, and, in these beds, fissures and cavities studded with the most beautiful crystals, which required, perhaps, centuries for their formation. These native crystals have a very different form from those prepared by fusion. They appear like two pointed four-sided pyramids, applied base to base (Fig. 53); such a form is called an acute octahedron, because contained under eight acute triangles. Thus sulphur, like carbon in the diamond and graphite, assumes two different crystalline forms; it is *dimorphous.*

Fig. 53.



Experiment 3.—Although sulphur is insoluble in *water,* many liquids are known which easily dissolve it; the one most convenient for this purpose is carbon disulphide. Pour a little of this liquid into a small clean bottle, and add a few pieces of sulphur. Close the bottle, and shake it at intervals for an hour or two. Then filter the liquid from the undissolved sulphur into a small basin, and allow it to evaporate; as it does so *crystals* of sulphur of the form shown in Fig. 53 will appear. Their shape differs from those obtained by fusion, but is similar to the native crystals.

Experiment 4.—Put about two ounces of fragments of roll-sulphur into a dry Florence flask, and place the latter on a retort stand. Heat it gently at first, but afterwards over a Bunsen's gas flame. (Fig. 7, page 27.) The sulphur will first melt to a thin yellowish liquid. When it is all melted, and when the heat is increased, the character of the liquefied sulphur is remarkably altered; it loses its mobility, and becomes thick and pasty, and dark in colour like treacle; so much so that the flask may be quite inverted without danger of the contents running out. As the heat continues, this condition is likewise changed, and soon the sulphur becomes again a *thin* liquid, which retains, however, its dark colour, and thus it continues until the boiling-point is reached. If the sulphur be thrown into water while in this condition, a substance entirely different in appearance from ordinary sulphur is obtained in the form of a reddish-coloured, soft, elastic mass, which serves admirably as a medium for taking impressions of seals, medals, &c. But this change is not a permanent one; in a short time the mass becomes hard and brittle, and is soon reconverted into ordinary sulphur. This plastic form of sulphur is said to be *amorphous*, a term applied to other bodies having no regular form, such as gum, pitch, glue, &c.

Experiment 5.—If, instead of pouring out the sulphur as in the last experiment, the heat be further applied, the sulphur will soon give out a rich dark-red vapour, which gradually occupies the whole of the upper part of the flask: thus sulphur is *volatile*, and may, like water, assume all the three states of aggregation (solid, fluid, and aeriform). Solid sulphur has a specific gravity of 2; that, is twice as heavy as water. Sulphur vapour has a specific gravity of 96 ($H = 1$), but at a very high temperature (1000° C) it trebles its volume, and has therefore a specific gravity of only 32. The formula for the ordinary vapour is, therefore, S_6, and for the vapour at high temperatures, S_2. Within the flask the vapour of the sulphur is transparent, and has a reddish-brown colour; but after escaping it appears as a yellowish smoke, being condensed by the cold air into a dust of solid sulphur. If this vapour be conducted into a glass jar, immersed in cold water, the sulphur condenses in it in the form of a soft yellow powder, known in commerce by the name of *flowers of sulphur*. The process by which a volatile substance is evaporated and condensed again into a solid is called *sublimation*. In distillation, the vapour is condensed into liquid (the distillate); in sublimation, into a solid (the sublimate).

The flowers of sulphur of commerce are obtained by vaporizing the impure *native* sulphur in iron or earthen retorts, and conducting the vapour into large chambers built of brick or stone, where it condenses as a crystalline powder, and lodges on the walls of the chambers. The interior of the chamber is sometimes allowed to become sufficiently hot to melt the deposited sulphur, which therefore runs down in the melted condition, and collects on the floor. While still liquid it is drawn off and solidified in wooden moulds, and so forms *roll-sulphur*, or *brimstone.*

Experiment 6.—Fill a test-tube half full of solution of caustic soda; add to it as much flowers of sulphur as can be taken up on the point of a knife, and boil the mixture for some time; a part of the sulphur will be dissolved, imparting to the liquid a yellowish-brown colour. The clear liquid is now decanted, diluted with water, and vinegar added to it; it will immediately assume a milky appearance, owing to the separation of the sulphur in the form of an exceedingly fine powder, which is so light that a considerable time must elapse before it will subside. Collect the powder on a filter, wash it with water, and dry it at a gentle heat. It is called *milk of sulphur*, or *precipitated sulphur*, and is sulphur in its finest state of subdivision. Precipitated sulphur has a pale yellowish tint, but on being melted it becomes distinctly yellow, owing to the union of the individual particles into a larger mass.

Experiment 7.—If sulphur be heated in a vessel with free access of air, as in an iron spoon, or be touched by some red-hot body, it *burns* with a blue flame; that is, it unites with the oxygen of the air, under the phenomenon of fire, and forms with the oxygen, as has been previously shown, a suffocating gas, sulphurous *anhydride* (SO_2).

This property which belongs to sulphur, of igniting at a very moderate heat, is the reason of its being so commonly used for all *kindling purposes.* By means of it, other bodies more difficult of combustion may be heated to the temperature at which they can continue to burn (matches, gunpowder, fireworks, &c.). The kindling of a simple coal-fire well illustrates how, by gradual transition from easily inflammable materials to those of more difficult ignition, the latter are finally brought to that degree of heat at which they will ignite and continue to burn.

Sulphur is an energetic element, and has, like oxygen, a powerful affinity for other elements.

Experiment 8.—Boil some sulphur in a flask, and expose a very thin copper-plate to the brownish vapour; the copper will glow vividly for some moments, lose its red colour and flexibility, become grey and brittle, and weigh one-fourth more than before. The newly-formed grey crystalline body is copper sulphide, CuS. Both elements have intimately combined, and in fixed proportions. The properties of the sulphur, as well as of the copper, have entirely disappeared. The great heat produced is a consequence of the chemical combination.

In a similar manner almost all other metals may be converted into sulphides. We find many, however, already formed in the earth, and miners call them glance, blende, or pyrites. The pyrites having the lustre of brass, and found in almost all coal, is sulphide of iron; red cinnabar is a sulphide of mercury, &c. The sulphide of copper, artificially prepared as above, occurs also as an ore, and is then called *copper pyrites.*

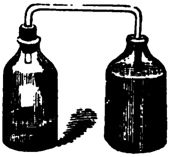
Experiment 9.—Mix three-fourths of an ounce of iron-filings, half an ounce of flowers of sulphur, and one fourth of an ounce of water, in a small vessel, and put it in a warm place; the mass becomes heated, the water evaporates, and in half an hour a black powder will be obtained, in which no particles of iron or of sulphur will be perceived; a chemical compound, *ferrous sulphide,* is formed. If the two substances be mixed together without water, no combination will take place, unless they be heated to redness; the water seems to effect the combination, by bringing the particles of the sulphur and iron into such close contact that they can attract each other. It is, as it were, the bridge by which one body passes over to the other.

Experiment 10.—Melt an ounce of sulphur in a covered earthen crucible, and then gradually throw in rather more than half an ounce of iron filings, closing the crucible after each addition. The iron will glow and combine with the sulphur just as the copper did in Experiment 8. The compound formed is the same as that obtained in Experiment 9, namely, *ferrous sulphide,* FeS. When larger quantities are required it is better to buy than to make it.

SULPHURETTED HYDROGEN, OR HYDROSULPHURIC ACID, H_2S.

Experiment 1.—Put half an ounce of ferrous sulphide (FeS) and half an ounce of diluted sulphuric acid into a two-ounce bottle, and quickly stop the bottle with a cork, to which a bent glass tube is adapted. Introduce the longer limb of the tube into a bottle filled with cold water. The atmospheric air contained in the flask and tube first passes over, followed by a very offensive gas, which dissolves in the water, to which it likewise imparts its fetid odour of rotten eggs. This gas is called *sulphuretted hydrogen.* The following change takes place:

Fig. 54.



$$FeS + H_2(SO_4) = Fe(SO_4) + H_2S.$$

When the disengagement of the gas ceases, add some diluted sulphuric acid, that the gas may again be generated. The water is known to be saturated with the gas when, on shaking the bottle, the finger by which the opening is closed is no longer sucked in, or, more correctly speaking, pressed in. One measure of water contains about two and a half measures of gas in a saturated solution. It is put up in small well-stoppered bottles. If the air be admitted, the solution becomes turbid, owing to the oxygen of the air uniting with the hydrogen of the sulphuretted hydrogen, forming water, and the consequent liberation of the sulphur as a fine powder.

If, during the evolution of the gas, the bottle of water be removed, the gas issuing from the tube can be ignited by a match; it burns with a blue flame, and its nauseous odour is no longer perceptible, but is replaced by the well-known odour of burning sulphur. Both constituents unite with the oxygen of the air, the sulphur forming sulphurous anhydride, and the hydrogen, water.

The inhalation of sulphuretted hydrogen is detrimental to health; hence precautions should be taken to avoid it. When experimenting with it, it is best to do so where there is a free circulation of air. A cloth moistened with a little alcohol, and held before the mouth, is likewise a good protection.

Sulphuretted hydrogen turns blue litmus-paper red; it also combines with bases, and hence it may be regarded as an acid.

Experiment 2.—Drop some sulphuretted hydrogen water upon a bright silver or copper coin, and upon a piece of lead and iron. The first three metals tarnish quickly, and finally become black; they combine with the sulphur, forming dark *metallic sulphides*, whilst the hydrogen escapes; the iron, on the contrary, undergoes no change.

Experiment 3.—Put into one test-tube a small portion of litharge, into another some ignited iron-rust, and pour upon them liquid hydrosulphuric acid. The yellow litharge, oxide of lead, becomes immediately black, an exchange of elements takes place, the hydrosulphuric acid gives its sulphur to the lead of the litharge, and receives in return the oxygen of the latter. Accordingly, lead sulphide and water are formed, and the offensive odour disappears. In the vessel containing the iron-rust, neither the colour nor the smell is changed—a proof that no chemical change has taken place.

Experiment 4.—Repeat the same experiment with a small crystal of sugar of lead instead of the litharge, and some green vitriol instead of the iron-rust, together with a few drops of vinegar, these salts having been previously dissolved in a large quantity of water; the result will be the same as in the former experiment. Sugar of lead is lead acetate; the salt of lead is converted into sulphide, which subsides sooner or later as a black precipitate. When this solution is extremely diluted it is only rendered brown. Acetic acid is set free, and remains in solution.

Most of the metallic sulphides are *insoluble* in water; hence sulphuretted hydrogen is peculiarly adapted for precipitating metals from their solution, so that they can be separated and collected by filtration. If sulphuretted hydrogen be passed through a solution of acetate or sulphate of copper, sulphide of copper will be precipitated, and can be separated by filtration from the acid. All the sulphides do not possess a black colour; sulphide of antimony has an orange-red colour; sulphide of arsenic a yellow, and sulphide of zinc a white colour. On this is partly based the application of sulphuretted hydrogen as a test; that is, as a means of detecting many metals. Wine containing lead is blackened by hydrosulphuric acid.

Many metals are precipitated from their acidified solutions by the addition merely of sulphuretted hydrogen, as sulphides; for example, copper, silver, gold, lead, mercury, tin, antimony, and arsenic; and others are not precipitated until a base is added; for example, iron, zinc, manganese, cobalt, and nickel. Sulphuretted hydrogen may accordingly be used to separate one metal from another; it is, therefore, an important means of separation in analytical chemistry.

Sulphuretted hydrogen has, as already mentioned, the formula H_2S, which indicates that it is composed of two atoms of hydrogen and one of sulphur, and the similarity of this formula to that of water, H_2O, is apparent. Lead paper is used for the detection of sulphuretted hydrogen, by which it is coloured brown or black. It is made by passing strips of paper through a weak solution of sugar of lead in water.

Finally, it remains to be stated that this gas occurs also in some *mineral waters*, as may be recognised by the smell and taste. Many of these springs—for instance, the celebrated springs of Harrogate and Aix-la-Chapelle—are resorted to by invalids, and are called *sulphur springs*. A rotten wooden water-pipe may convert an otherwise potable water, if it should contain calcium sulphide, into a nauseous sulphuretted water; but by clearing out the well, and laying down new pipes, the water may be rendered completely odourless and good.

COMPOUNDS OF SULPHUR AND OXYGEN.

Sulphur combines with oxygen in two proportions, forming the oxides SO_2 and SO_3; they both unite with water to form acids. SO_2 is hence called sulphurous *anhydride*, and SO_3 sulphuric anhydride. The latter substance can only be prepared indirectly and with difficulty, whereas sulphurous anhydride is formed by the direct union of its constituents whenever sulphur is burnt in air or oxygen, as has been already mentioned (page 124).

Sulphurous Anhydride, SO_2, *and Sulphurous Acid*, H_2SO_3.—Sulphurous anhydride is a colourless gas of a pungent suffocating odour, well known as the smell of sulphur. One of its most characteristic properties is its power of bleaching animal and vegetable colours.

Experiment 1.—Suspend a red flower, such as a rose or peony, in an inverted glass jar, or other similar vessel, and introduce a piece of ignited sulphur. It will be found that the colour of the flower is slowly changed to white. Sulphurous anhydride, unlike chlorine, does not actually *destroy* the colouring matter, for by dipping the flower into dilute sulphuric acid the original colour is slowly restored. Straw, wool, silk, &c., are very commonly bleached with sulphurous anhydride, but the result of its action on these materials is far from permanent; for on exposing a bleached straw hat in the sunshine it again becomes yellow, or "sun-burnt," as the effect is popularly termed; and on washing whitened flannel it resumes its original appearance.

Sulphurous anhydride can be formed in a variety of ways; one of the most convenient, when a quantity is required in a tolerable state of purity, is the following:

Experiment 2.—Select a thin-bottomed, well-made Florence flask, fit it with a delivery tube, and place in it half-an-ounce of thin copper, sheet or turnings, and two ounces of strong sulphuric acid. Apply a gentle but gradually increasing heat, and after a time sulphurous anhydride will be evolved, according to the equation:

$$Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2.$$

Sulphurous anhydride is more than twice as heavy as air, so it may be collected in dry bottles by *displacement*; that is, by allowing the gas to pass to the bottom of a dry bottle. It will gradually displace the lighter air, just as water does when poured into a bottle, and as it is very soluble in water this mode is the most convenient.

Experiment 3.—If a lighted taper is held over burning sulphur, or introduced into a vessel of sulphurous anhydride gas, it will be *extinguished*. It may be now readily explained how chimneys on fire are extinguished by scattering sulphur on the coals beneath; the sulphurous acid gas ascends in the chimney, and expels the atmospheric air present in it; the glowing soot is thereby deprived of the free oxygen, and is extinguished.

Experiment 4.—Pour a little blue litmus solution into a bottle of sulphurous anhydride; it will be reddened, showing the acid character of the gas.

Experiment 5.—Dip the end of the delivery tube from which the gas is issuing into a little water; the gas is absorbed in large quantities and imparts to the water its characteristic taste and smell. One measure of water dissolves about forty measures of sulphurous anhydride. The solution obtained is one of sulphurous *acid:*

$$SO_2 + H_2O = H_2SO_3.$$

The salts corresponding to sulphurous acid are called *sulphites:*

H_2SO_3, Sulphurous acid.
Na_2SO_3, Sodium sulphite.
$Ca''SO_3$, Calcium sulphite.

Experiment 6.—When the gas ceases to be absorbed by the water, substitute for the latter a solution of sodium carbonate; this likewise absorbs the gas and forms with it sodium sulphite, while the carbonic acid is liberated with effervescence. If a sulphite be treated with a strong acid it is decomposed and sulphurous anhydride is evolved.

Experiment 7.—Place a few grains of sodium sulphite in a test-tube, pour upon it a little slightly diluted sulphuric acid, and gently warm. Sulphurous anhydride is given off and is readily identified by its odour:

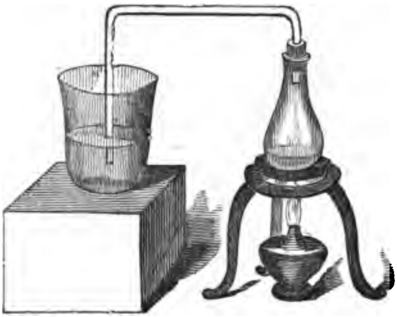
$$Na_2SO_3 + H_2SO_4 = Na_2SO_4 + H_2O + SO_2.$$

Experiment 8.—When the residue in the flask, in which the sulphurous anhydride was generated, has become cold, add water to it, and heat it gently to boiling until the residual mass is dissolved. The solution is dark and turbid from impurities contained in the metal, but after filtering it is of a beautiful blue colour, and transparent. If allowed to cool slowly, blue crystals of copper sulphate, $Cu''SO_4 5H_2O$ (blue vitriol), of considerable size will be formed. Silver and mercury comport themselves like copper and may be used instead, but they are not as economical.

Sulphuric anhydride, SO_3, *and Sulphuric Acid*, $H_2(SO_4)$.

Sulphuric Anhydride.—Experiment 9.—Pour into a small flask, placed in a sand-bath over a tripod, half an ounce of *fuming*, or Nordhausen sulphuric acid, and heat it gently till it boils moderately. Conduct the vapour through a tolerably wide glass tube into an empty flask, not shown in the cut, which is placed in a vessel filled with cold water. In summer time the water may easily be made colder by adding a few tea-spoonfuls of powdered saltpetre. If the vapour be suffered to escape into the air, it appears in thick white fumes, having a pungent acid smell; but if conducted into the flask, it is condensed into a glistening white, solid mass. This is sulphuric anhydride. The distillation is stopped as soon as the boiling ceases, and the glass tube becomes too hot for the hand to bear. What remains in the flask no longer fumes; it has become common sulphuric acid. To cause this to boil, you must apply a ten-times stronger heat than before, for it does not begin to boil till above 572° F. (300° C.), while the anhydride boils even at a little above 86° F. (30° C.) This is the reason why the boiling ceases when the latter has escaped.

Fig. 55.



Experiment 10.—Take out some of the anhydride by means of a glass rod, and introduce it into a dry test-tube; it will fume violently, and after a time become fluid; *that is, it attracts water* from the air, and is thereby converted into Nordhausen sulphuric acid. On longer standing, it absorbs still more water, and ceases to fume; it thus becomes common sulphuric acid. By evaporation this water cannot again be removed.

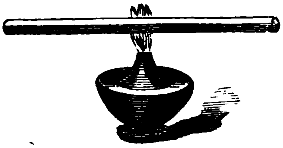
Experiment 11.—If anhydrous sulphuric acid be thrown into water, it is dissolved with a hissing noise and the *violent evolution of heat.*

Experiment 12.—It is likewise dissolved by common sulphuric acid, converting this into the fuming acid. Fuming sulphuric acid may therefore be regarded as a compound of the anhydride with ordinary sulphuric acid, H_2SO_4,SO_3.

Fuming, or Nordhausen sulphuric acid is obtained by the distillation of green vitriol, or ferrous sulphate.

Experiment 13.—Put a crystal of green vitriol into a hard glass tube, and heat it; aqueous vapour escapes, and the green crystal becomes white (anhydrous). On further heating the white colour passes into reddish-brown, and sulphurous and sulphuric anhydrides are evolved, while the iron remains as ferric oxide:

Fig. 56.



$$2FeSO_4 = SO_2 + SO_3 + Fe_2O_3.$$

In preparing the fuming acid on a large scale, earthen retorts are used, and the anhydride is conducted into common sulphuric acid, which dissolves the sulphuric anhydride, and is thereby converted into the fuming acid, while the sulphurous anhydride escapes. It is a thick liquid, like oil, and is called Nordhausen sulphuric acid, because this city supplied Germany with it for centuries. It has the specific gravity of 1·9.

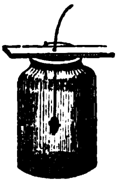
If this acid be exposed to the air, the anhydride in it evaporates, and unites with the watery vapour contained in the air; accordingly, common sulphuric acid is formed, which, being less volatile, condenses in the cold air, forming white vapours, just as steam does. Consequently the fumes of this acid consist of the vapour of common sulphuric acid.

As long as the process of manufacturing sulphuric acid from green vitriol was the only one known, it was a very expensive article. A hundredweight is now obtained for the same sum that was formerly paid for two pounds.

Common sulphuric acid, H_2SO_4.—Sulphur by burning is converted into sulphurous anhydride, SO_2, and this combined with water gives sulphurous acid, H_2SO_3. To convert this into sulphuric acid, H_2SO_4, an additional atom of oxygen is necessary, and this additional oxygen is generally supplied at the expense of nitric acid.

Experiment 14.—Fasten some pieces of sulphur to an iron wire, inflame, and hold them in a capacious bottle containing a little water, until the blue sulphur flame is extinguished; the bottle becomes filled with a white smoke, which is recognised by its odour to be sulphurous acid. If you now introduce a shaving moistened with nitric acid into the vessel, reddish-yellow fumes will immediately form around the wood, gradually filling the whole bottle. These fumes are nitric peroxide, and their evolution indicates that the sulphurous acid has withdrawn oxygen from the nitric acid, and has been oxidized and converted into sulphuric acid. After some time the bottle becomes clear again, because the vapour of the sulphuric acid formed sinks to the bottom, and dissolves in the water, and we can now again burn sulphur in the bottle. If we repeat this operation several times, we can soon prepare a few ounces of diluted sulphuric acid.

Fig. 57.



Experiment 15.—Add some drops of a solution of barium chloride to a portion of the acid liquid just obtained; a copious white precipitate is formed, which disappears neither by boiling, nor by the addition of water or nitric acid. This precipitate is *barium sulphate*, a salt quite insoluble in water and acids. Add one drop of the diluted acid to a wine-glassful of water, and add to this a little more of the solution of barium chloride; even at this great dilution a perceptible cloudiness will be produced. A solution of barium chloride, or nitrate, is the most certain test for detecting sulphuric acid and sulphates.

The manufacture of common sulphuric acid on a large scale is conducted on the same principle as in the last experiment but one. The sulphur is burnt on a sort of hearth, and the sulphurous anhydride thus formed is carried by the draught of air into an immense leaden chamber, the atmosphere of which is kept moist by jets of steam. The sulphurous anhydride thus becomes sulphurous acid. A little nitric acid (or saltpetre and sulphuric acid, which yield nitric acid) is placed in a basin, supported by a tripod above the burning sulphur, and the heat drives it in vapour into the chamber. In the leaden chamber this nitric acid immediately oxidizes a portion of sulphurous acid into sulphuric acid, and is itself reduced to nitric oxide, NO:

$$2HNO_3 + 3H_2SO_3 = H_2O + 2NO + 3H_2SO_4.$$

Now we shall hereafter learn that nitric oxide is a colourless gas which, on contact with air, takes up oxygen and forms the brown gas, nitric peroxide, NO_2, $NO + O = NO_2$. This change immediately takes place in the leaden chamber, and as fast as the NO_2 is formed it is again reduced to NO by another portion of sulphurous acid:

$$NO_2 + H_2SO_3 = NO + H_2SO_4.$$

Again the NO takes oxygen from the air of the chamber, and becomes NO_2, and again a new portion of sulphurous acid is oxidized into sulphuric acid, so that the process, after the first stage, is continuous. The nitric oxide acts as a *carrier*, taking oxygen from the air, and giving it out again to the sulphurous acid, which is unable to take oxygen direct from the air (except very slowly), so that theoretically there is no limit to the quantity of sulphurous acid which may be oxidized by a given weight of nitric acid. As fast as the sulphuric acid forms, it condenses and settles down on the floor of the leaden chamber. From time to time it is drawn off and concentrated by evaporation, first in leaden, and afterwards in glass or platinum vessels. The acid thus obtained has almost exactly the composition of true sulphuric acid, H_2SO_4.

The above account of the sulphuric acid manufacture must only be regarded as an approximation to the truth, for the process is in fact extremely complex. If the quantity of steam be deficient, white crystals of complex composition are apt to form. On the addition of water, however, they are decomposed, and sulphuric acid formed from them.

Experiment 16.—Let some sulphuric acid remain in an open flask exposed to the air; it will increase every day in weight, *for it very eagerly attracts water from the air.* After standing for some months in a damp place it will become two or three times heavier than before. Some substances, especially gases, are dried by means of sulphuric acid.

Experiment 17.—A piece of wood introduced into sulphuric acid becomes black, and is charred just as when it is exposed to the flame of a candle. The sulphuric acid seizes upon its hydrogen and oxygen, which combine to form water, and the carbon is left behind. Wood may be charred in this way, in order to protect it from decay in moist situations. In the refining of lamp-oil, the mucilage of the oil is charred by sulphuric acid. *Sulphuric acid chars and destroys many vegetable and animal substances.* If figures are inscribed on paper with very dilute sulphuric acid, and the paper warmed, the figures will come out black. As the acid becomes concentrated, it acts upon and chars the paper.

Experiment 18.—Pour a drop of oil of vitriol upon paper; decomposition takes place slowly, but it will take place instantaneously if a drop of water is added, because *water and sulphuric acid unite together with the evolution of strong heat.* For this reason, when sulphuric acid comes in contact with the skin, it should first be wiped off with dry paper or cloth, and then be immediately washed with a great quantity of water. If 50 measures of sulphuric acid are mixed with 50 measures of water, we do not obtain 100 measures, but only 97 measures, of liquid; consequently, a contraction or condensation has occurred, which condensation is itself attended with the liberation of heat.

Experiment 19.—Pulverise a small quantity of *indigo*, and form a thin paste of it with *fuming sulphuric acid.* After a few days add to it some water, and you obtain a deep blue liquid—solution of indigo. With this solution wool may be dyed of a fine blue colour (Saxon-blue). Common sulphuric acid dissolves indigo only imperfectly. Indigo, although a vegetable substance, is not carbonised by sulphuric acid, thus forming an exception to the general rule.

Experiment 20.—To half an ounce of copper scales, such as fall off at the coppersmiths', add two ounces of water, and then gradually two-thirds of an ounce of sulphuric acid, and put it in a warm place; you obtain a blue solution, from which afterwards blue oblique, rhomboidal crystals will be deposited. The edges of these crystals are usually obtuse, giving to the narrow sides a roof-like appearance. Copper scales consist of copper oxide; and they combine with the acid, forming copper sulphate (blue vitriol), a *soluble* salt. The following metallic sulphates are commonly called vitriols: ferrous sulphate (green vitriol); zinc sulphate (white vitriol); and copper sulphate (blue vitriol). Copper or iron vessels are cleaned more rapidly and made brighter by water to which some sulphuric acid has been added, than by simple water alone, because the oxide, which tarnished the vessels, is dissolved by the acid.

Experiment 21.—Put a small iron nail into a test tube, and drench it with twenty drops of common sulphuric acid; it is not acted upon. But if you add a little water, about four or five times more than the acid, a brisk effervescence will ensue, and the iron will be dissolved. The strong acid may be heated to boiling in iron vessels without acting upon them, which is by no means the case with the diluted acid. In this experiment, hydrogen escapes, and iron sulphate remains in solution. Zinc behaves in a similar manner. Consequently, the *diluted acid* must be employed for dissolving such metals. But there are also metals which dissolve only in the stronger acid, with the aid of heat; for instance, copper, silver, &c. This has been treated of under sulphurous acid.

SELENIUM, Se = 79·5, AND TELLURIUM, Te = 129.

These elements are very scarce and unimportant; they occur combined with certain metals in some rare minerals. They both resemble sulphur in their properties, but partake more of the character of metals in appearance. Their compounds with hydrogen and oxygen correspond perfectly to those of sulphur.

CHAPTER III.

NON-METALLIC TRIADS.

(Nitrogen, Phosphorus, Boron.)

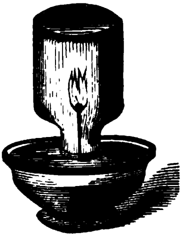
NITROGEN.

Symbol, N = 14. Formula, N_2.

In the free state as gas nitrogen constitutes nearly four-fifths of the atmosphere, the remaining one-fifth being oxygen. In a state of combination it is an important constituent of organic structures. It is also found in the mineral kingdom, as in saltpetre or nitre, a salt occurring as an incrustation on the earth in India and elsewhere; from this salt the name *nitrogen* (generator of nitre) is derived. The atomic weight of nitrogen is 14, and the molecule consists of two atoms. Hence the specific gravity of the gas is 14 (p. 72).

The most convenient and simple method for the preparation of nitrogen consists in removing the oxygen from a confined portion of air by the ordinary process of combustion.

Fig. 58.



Experiment 1.—Fasten a piece of sponge to a wire, drop some alcohol upon it, and hold the wire in a vessel containing water so that the sponge may be some inches above the water. Then kindle the spirit, and immediately place an empty bottle over it, so that the mouth of it may dip into the water; the flame will soon cease burning, and some of the water will rise into the bottle, in proportion to the amount of air disappearing during the combustion. The consumed air was oxygen, which united with the constituents of the alcohol. Close the bottle tightly with the finger, shake it briskly, and again open it below the water, when a little more water will enter. The air which remains in the bottle is *nitrogen;* it is sometimes called *azote* (*a*, privative, and ζωή, *life*), from its inability to support respiration.

Other combustibles may be used for the purpose of abstracting the oxygen from air, such as sulphur, charcoal, &c., but phosphorus is the best; for the products of combustion of the other substances, being gaseous, are not easily got rid of, and consequently contaminate the nitrogen, whereas the phosphorus compound produced is solid, and moreover very soluble in water. Moreover phosphorus burns till all the oxygen is removed, which is not the case with many combustibles.

Experiment 2.—Place a small piece of dry phosphorus in a little basin floating in water. Ignite the phosphorus and place over it a bell-jar or the mouth of a wide-necked bottle (Fig. 58). The phosphorus will burn until the whole of the oxygen is consumed. Dense white clouds of phosphoric anhydride (P_2O_5) are formed and fill the jar, but these after a time are precipitated in snowy flakes, and are rapidly dissolved by the water. The water will rise in the bottle, partly from the removal of oxygen and partly from some air having been expelled at the beginning of the experiment, in consequence of the expansion produced by the heat (p. 34). When the white fumes have entirely disappeared the bottle may be shaken so as to cause the basin to fill and sink, and the nitrogen transferred in the pneumatic trough to a series of smaller bottles.

Nitrogen is incombustible, except at a very high temperature. It will not support combustion, for a lighted taper immersed in a jar of it is extinguished. It has neither colour, smell, nor taste, and in a chemical point of view it must be regarded as a very inert body, since it does not combine directly with any other substance except with great difficulty.

Nitrogen gas is erroneously called *azote*, as we are continually breathing it without perceiving any injurious effects from it; it stops respiration only when it contains no oxygen, and because it contains none. The human body is so constructed, that it will not thrive on substances intended as nourishment if they are presented to it in their purest form. Strong alcohol acts as a poison, but when diluted with four or five times its quantity of water, as in wine, it is invigorating. Even the respiration of oxygen would soon destroy life, were it not diluted with four times its volume of nitrogen, as in atmospheric air.

Besides oxygen and nitrogen, air contains *vapour of water* and *carbonic anhydride.* The presence of the former is rendered obvious by the fall of rain, snow, dew, &c.,; and that of carbonic anhydride can easily be shown by letting limewater remain exposed to the air, or by shaking it in a flask containing air. This occasions a cloudiness in the liquid. If you ask, What is the source of this carbonic anhydride? the reply is, It is formed wherever substances are burning, wherever men and animals are breathing, and wherever decay and putrefaction are taking place. (Page 126.)

The following is a close approximation to the average composition of the air:

	By Volume.	By Weight.
Nitrogen	79	77
Oxygen	21	23
	100	100

The carbonic anhydride amounts in pure country air to about three volumes in 10,000. In towns it is more abundant.

The air is not a *compound,* but only a *mixture.* It varies slightly in composition, which compounds never do (page 63), and its properties are intermediate between those of its constituents, whereas those of a compound are entirely different.

In crowded rooms, and other confined places, the air becomes deteriorated; that is, poorer in oxygen and richer in carbonic anhydride.

That the air also contains other foreign ingredients is not strange, since it is the constant receptacle of volatile substances and dust. The air coming from the Spice Islands, even at the distance of eight or ten miles, is impregnated with the odour of cinnamon and cloves. The dust contained in the air can be discerned in the sunbeam, &c. Minute quantities of ammonia and nitric acid are also found in the air. The presence of ozone has already been noticed. The specific gravity of dry air, at standard pressure and temperature, is 14·47, that is, it is about 14½ times heavier than hydrogen.

AMMONIA, NH_3.

This important gas can be prepared in minute quantity by the direct union of its elements. It is also frequently formed during the putrefaction of organic bodies, but its most convenient source is in the "ammoniacal liquor" of gas-works. From this a white substance called *sal-ammoniac* —the ammonium chloride of chemists—is manufactured.

Experiment 1.—Mix about half an ounce of powdered sal-ammoniac with about one ounce of quick lime. A pungent odour of hartshorn is perceived. Introduce the mixture into a dry Florence flask fitted with a cork and a short tube which communicates by an india-rubber tube with a long straight tube. Apply a gentle heat, and ammonia gas will come off in abundance. It is very soluble in water, but being lighter than air it may be collected by *upward displacement*, that is, by holding a bottle mouth downwards over the tube and allowing the light gas to displace the heavier air. The reaction by which the gas is prepared is as follows:—

$$\underset{\text{Ammonium Chloride.}}{2NH_4Cl} + \underset{\text{Lime.}}{CaO} = \underset{\text{Ammonia.}}{2NH_3} + \underset{\text{Calcium Chloride.}}{CaCl_2} + H_2O.$$

Calcium chloride remains in the flask, together with the excess of lime.

Experiment 2.—Plunge a taper into a bottle of ammonia held with its mouth downwards. The taper will be extinguished and will not ignite the gas.

Experiment 3.—We must not too hastily conclude from the preceding experiment that ammonia cannot be made to burn. Hold the end of the tube from which the gas is issuing in a flame. You will perceive that the gas burns with a greenish flame *as long as the jet is heated*, but that it goes out when the lamp is removed. *Ammonia will burn when strongly heated.* During its combustion its hydrogen is oxidized to water, and its nitrogen escapes free.

Experiment 4.—Remove the stopper downwards from a bottle of ammonia, and *afterwards* immerse the mouth of the bottle in water. The water will rise rapidly, and if the bottle was quite full of gas, will fill it. If a piece of red litmus-paper is dipped in the water of the bottle its blue colour will return. *Ammonia is soluble in water, and its solution is a base.*

One volume of water at freezing point will dissolve 1149, and at ordinary temperatures about 700 volumes of the gas, so that it is even more soluble than hydrochloric acid. A concentrated solution of it may be prepared like that of hydrochloric acid (page 113). It is sold under the name of solution of ammonia, or "liquor ammoniæ," and has a specific gravity of 0·88. Weaker solutions are sometimes called "hartshorn," or "spirits of hartshorn," because they were formerly obtained by the distillation of horn.

Experiment 5.—Ammonia gas, like hydrochloric acid, can easily be prepared by gently heating its concentrated solution. By prolonged boiling all trace of ammonia may be removed from water.

Experiment 6.—The direct combination of ammonia and hydrochloric acid gases has already been noticed (page 13). The experiment may be repeated with equal volumes of the pure gases. Sal-ammoniac or ammonium chloride is formed:

$$NH_3 + HCl = NH_4Cl.$$

Experiment 7.—Carefully neutralize a dilute solution of ammonia with nitric acid, so that it has no effect on either blue or red litmus. The solution will contain a *salt* called ammonium nitrate, which may be obtained in the dry state by evaporation:

$$NH_3 + HNO_3 = NH_4NO_3.$$

Ammonium compounds.—The compounds of ammonia are so similar in many of their properties to those of potassium and sodium, that they are generally supposed to be analogous in structure, and to contain a sort of compound metal called *ammonium*, NH_4, which has recently been isolated. When the gas dissolves in water it is assumed that the hydrate of this *quasi* metal is formed: $NH_3 + H_2O = NH_4HO$. The theory cannot be proved, but it is probable and useful. The following table shows the analogy which on this view exists between potassium and ammonium salts:—

K HO	Potassium	Hydrate.	NH_4 HO	Ammonium	Hydrate (?).
K Cl	„	Chloride.	NH_4 Cl	„	Chloride.
K NO_3	„	Nitrate.	NH_4 NO_3	„	Nitrate.
K_2SO_4	„	Sulphate.	$(NH_4)_2$ SO_4	„	Sulphate.

It will be seen that to represent the composition of any ammonium salt we need only take the formula of the corresponding potassium or sodium compound and write NH_4 instead of K or Na. It must be remarked that although solution of ammonia *may* contain the hydrate NH_4HO, no proof that such is the case has been obtained.

Experiment 8.—Put about an ounce of mercury into a mortar and press into it with the pestle a few fragments of clean dry sodium. The two metals will combine with a flash of light, and a semi-fluid mass called *sodium amalgam* will be obtained. When cold, throw this into a tumbler about one-third full of a cold saturated solution of sal-ammoniac. The amalgam will immediately begin to swell, and will soon assume such an enormous volume as to float in the liquid as a metallic mass, which feels like butter. From the moment of its formation, however, this curious substance begins to change into ammonia, hydrogen and mercury. It is called *ammonium amalgam*, and it is believed to be formed by the temporary combination of the ammonium, which has lost its chlorine, with mercury. The following formula, in which *n* denotes an unknown number of atoms, explains its formation:

$$NH_4Cl + NaHg_n = NH_4Hg_n + NaCl.$$

By washing with water the mercury is afterwards recovered without loss.

The specific gravity of ammonia gas is 8·5, a little more than half that of air (14·47). By a pressure seven times that of the atmosphere the gas can be condensed into a liquid, and if cooled to $-75°$ C. ($-103°$ F.) it freezes to a transparent, ice-like solid.

It must be borne in mind that the name ammonia is often given to the solution, which *should* be called ammonium hydrate.

COMPOUNDS OF NITROGEN AND OXYGEN.

	Oxides.	Corresponding Acids.
N_2O	Nitrous oxide.	
N_2O_2 or NO	Nitric oxide.	
N_2O_3	Nitrous anhydride.	HNO_2 Nitrous acid.
N_2O_4 or NO_2	Nitric peroxide.	
N_2O_5	Nitric anhydride.	HNO_3 Nitric acid.

Of these compounds nitric acid is undoubtedly the most important, and as, moreover, it is the source from whence the other oxides are obtained, it claims our first attention.

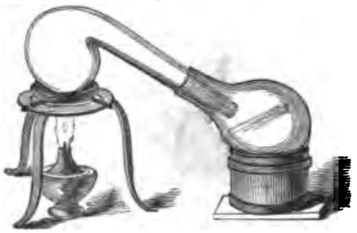
The oxide N_2O_5, nitric anhydride, corresponding to nitric acid, is prepared with difficulty. It is a white, crystalline solid, which combines eagerly with water:

$$N_2O_5 + H_2O = 2HNO_3.$$

Nitric Acid or Aquafortis, HNO_3.

Experiment 1.—Introduce into a small retort half an ounce of powdered saltpetre and half an ounce of common sulphuric acid, and let the retort stand erect for some time, in order that as much as possible of the sulphuric acid remaining in the neck may flow down into the retort. Then surround the latter with sand contained in an iron basin, adapt to the beak of it a receiver, wrap round the joint some strips of moistened blotting-paper, and heat gently. In a short time a yellowish fuming fluid passes over into the receiver, which is placed in a vessel filled with water, and must frequently be sprinkled with cold water; this fluid is heavier than water, and is called *nitric acid.*

Fig. 59.



Saltpetre is potassium nitrate, KNO_3. When it is acted upon by sulphuric acid the metal and part of the hydrogen change places; we get nitric acid and a salt, called hydrogen potassium sulphate, $HKSO_4$, in which half the hydrogen of the acid is replaced by potassium :

$$KNO_3 + H_2SO_4 = HNO_3 + HKSO_4.$$

At a higher temperature the hydrogen potassium sulphate will decompose another molecule of potassium nitrate:

$$KNO_3 + HKSO_4 = HNO_3 + K_2SO_4;$$

but on the small scale it is better not to push the action so far or the retort will be apt to break. K_2SO_4 is called neutral, or di-potassium sulphate.

By using perfectly dry potassium nitrate and concentrated sulphuric acid, true nitric acid, HNO_3, can be obtained as a yellowish liquid, which fumes strongly when exposed to the air. It is hence called *fuming* nitric acid, and is the strongest that can be prepared. It is one and a half times heavier than water, its specific gravity being 1·52. A weaker kind is commonly met with in commerce as ordinary nitric acid, or *aquafortis*. It consists of 60 parts of true nitric acid and 40 of water, and has a specific gravity of 1·42. This acid is colourless when pure, but usually possesses a yellowish tint. When very strong or very weak nitric acid is heated, acid of this strength is obtained. It can be distilled unchanged.

Experiment 2.—A drop of nitric acid is sufficient to acidify several spoonfuls of water, and even at a greater dilution it will redden blue litmus-paper; nitric acid is accordingly distinctly characterised as an *acid*.

Experiment 3.—If lead be heated for a long time in the air it abstracts oxygen from it, and becomes converted into a reddish-yellow powder, called lead oxide, or litharge. Take up a small portion of this litharge on the point of a knife, put it into a test-tube, and add some dilute nitric acid. The greater part will be dissolved by gentle heating. Filter the solution while warm, and put it in a cold place; a salt will be deposited from it in white brilliant crystals; this is *lead nitrate*. This shows that lead oxide is a basic oxide, as it combines with acids forming salts.

Nitric acid dissolves most of the metallic oxides, and forms with them salts, all of which are soluble in water. For this reason, nitric acid is often used for cleaning metals, for instance, copper and brass instruments, which, during the process of annealing, soldering, &c., have become covered with a coating of oxide.

Experiment 4.—Pour over some shot, common nitric acid, slightly diluted with water; a solution is also effected in this instance, but it is accompanied by the evolution of a yellowish-red vapour of a suffocating smell. This vapour is nitric peroxide. Part of the nitric acid is decomposed, while another part of it combines with the lead, and forms the same salt, as in the former experiment. This likewise crystallises from its solution, if it is evaporated until a film forms on its surface.

In this case the lead is apparently dissolved, but it is obvious that this is quite a different kind of solution from that of common salt or sugar in water. The salt and sugar are unchanged in the solution, while the lead is not contained in the liquid as a metal, but as a salt, a nitrate. The same thing occurs with all other metals which are soluble in nitric acid; as, for example, with silver, mercury, copper, iron, &c. Gold is not dissolved by it; hence gold may be separated from silver by means of nitric acid.

Nitric acid readily parts with a portion of its oxygen. It is therefore a powerful *oxidizing agent.*

Experiment 5.—Place a few fragments of tin in a wine-glass, and pour over them a little nitric acid. A violent action is set up, and red suffocating fumes are copiously evolved. The tin is not dissolved and converted into nitrate as the lead was, but becomes a white insoluble powder, an oxide of tin called *meta-stannic acid.*

Experiment 6.—Some of the non-metallic elements, as well as of the metals, are oxidized by nitric acid; charcoal, on being boiled in it, becomes carbonic anhydride; sulphur, sulphuric acid; phosphorus, phosphoric acid, &c. In all these cases yellowish-red fumes are evolved.

Experiment 7.—Organic substances also, as wool, feathers, wood, indigo, &c., are oxidized and decomposed by heating them with nitric acid. This sort of decomposition may be regarded as combustion in the moist way. If substances of animal origin are allowed to remain for a short time only in contact with this acid, they will assume a yellow colour. In this manner wood may be stained, and silk may be dyed yellow; the hands and clothes are also stained yellow by nitric acid. Cotton undergoes a most remarkable change if soaked for a short time in the *strongest* nitric acid; it will then detonate and explode, like gunpowder, only far more violently (gun-cotton). Strong nitric acid is partially decomposed, and coloured yellow, by the rays of the sun.

If you colour some water blue in a test-tube with one drop of solution of sulphate of indigo, and add to it on boiling one drop of nitric acid, the blue colour will disappear. This reaction often serves for the detection of nitric acid.

Experiment 8.—The nitrates are easily decomposed. Having powdered some of the nitrate of lead, obtained in Experiment 3 or 4, throw it upon a red-hot coal; active combustion of the coal will ensue, at the expense of the nitrate, and globules of metallic lead will remain beyond.

Nitric acid is *monobasic;* that is, it contains but one atom of hydrogen which can be replaced by metals. Its salts are called nitrates, and it must be remembered that they all contain the monad radical, NO_3.

HNO_3 Nitric acid.
KNO_3 Potassium nitrate.
NH_4NO_3 Ammonium nitrate.
$Cu''(NO_3)_2$ Copper nitrate.
$Bi'''(NO_3)_3$ Bismuth nitrate.

The nitrates are less easily decomposed than the chlorates, but they resemble the latter salts in the readiness with which they part with their oxygen to combustibles. In consequence of this property, potassium nitrate is a constituent of some inflammable and explosive mixtures, such as gunpowder, fuses, coloured fires, &c.

Nitrous Oxide, N_2O.

Experiment 1.—Cautiously heat the ammonium nitrate, made by Experiment 7 (page 156), in a Florence flask to which a gas delivery tube is connected. The salt will gradually melt and soon effervesce briskly, as a gas—nitrous oxide—is evolved. When the air has been expelled from the flask the issuing gas may be collected at the pneumatic trough, which must be filled with warm water instead of cold, as the gas is soluble in the latter.

The change which occurs is simple. Heat decomposes ammonium nitrate NH_4NO_3 entirely into *water and nitrous oxide:*

$$NH_4NO_3 = 2H_2O + N_2O.$$

Nitrous oxide is seen to be a colourless gas. It can, however, be condensed to the liquid state by great pressure.

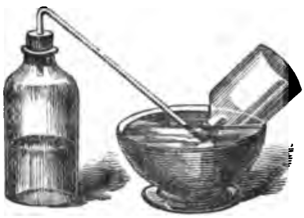
It is respirable in moderate quantity, and its effects when inhaled are peculiar. At first it produces a pleasurable kind of intoxication—whence its name of *laughing gas;* but after a time it causes complete insensibility to pain. It is now much used as a substitute for chloroform in minor surgical operations.

Nitrous oxide is a good supporter of combustion, its power in this respect being little inferior to that of oxygen. The experiments which were performed with the latter gas may be repeated with nitrous oxide, with little difference in their results.

Nitric Oxide, NO, *or* N_2O_2.

Experiment 1.—Pour over a few scraps of copper, placed in a wide-mouthed bottle, a little water, and then add by degrees some nitric acid, until a brisk effervescence ensues. This effervescence is caused by the evolution of nitric oxide, which must be collected in a jar of white glass over the pneumatic trough. Close the mouth of the jar under water; it seems to be empty, for the nitric oxide is colourless; but if the jar be opened, and air be carefully blown in, then the jar becomes filled from above downwards with yellowish-red vapours. The nitric oxide takes thereby from the air one atom of oxygen, and is converted into nitric peroxide. NO becomes NO_2. On account of this property, it has an important application in the preparation of common sulphuric acid (page 148).

Fig. 60.



The following is the somewhat complicated formula which describes the formation of nitric oxide:

$$\overset{\text{Nitric Acid.}}{8HNO_3} + 3Cu = \overset{\text{Copper Nitrate.}}{3Cu''(NO_3)_2} + 4H_2O + 2NO.$$

The copper nitrate remains as a blue solution in the bottle, and it may be obtained in crystals by evaporation. Nitric oxide supports combustion, although not with the readiness of nitrous oxide. If a piece of phosphorus be ignited and plunged into nitric oxide while feebly burning, it will be extinguished, but if the phosphorus burns briskly when immersed in the gas the combustion proceeds with energy.

Nitrous anhydride, N_2O_3, is a red gas, easily condensed to a blue liquid by cold. With water it forms the very unstable nitrous acid, HNO_2.

Nitric peroxide, NO_2, *or* N_2O_4, is the red gas formed when nitric oxide comes in contact with excess of oxygen. It is absorbed by water and converted into a mixture of nitrous and nitric acids:

$$2NO_2 + H_2O = HNO_2 + HNO_3.$$

PHOSPHORUS.

Symbol, $P = 31$. Formula for vapour, P_4.

Great care is required in experimenting with phosphorus that it does not take fire at an unseasonable moment, as it continues burning with the greatest violence, and might occasion dangerous wounds. It may catch fire even when lying upon blotting-paper, particularly in summer-time, or by the heat of the finger. Hence it must be kept, and also cut, under water. On being taken from the water, it should be held by a pair of forceps, or be stuck on the point of a knife. Prudence also would dictate to experiment with small quantities only at a time, and to have a vessel of water in readiness, in which it may be quenched in case it should catch fire.

Phosphorus, like sulphur, melts, boils, evaporates, and burns, but far more easily and rapidly. In winter it is brittle, in summer flexible as wax. When pure and freshly prepared it is colourless, transparent, and amorphous (non-crystalline), but after a time it becomes yellow, and coated over with a white crust.

Phosphorus is insoluble in water, but soluble in ether, carbon disulphide, and oils.

Phosphorus is an exceedingly violent *poison*, and is, for this reason, frequently employed for the extirpation of rats and mice. The rat paste, as it is called (phosphorus dough), is composed of one drachm of phosphorus, 8 ounces of hot water, and 8 ounces of flour.

Preparation of Phosphorus.—Phosphorus was formerly obtained from urine, but is now prepared from bone-ash. Bones consist chiefly of gelatin (or rather ossein) and calcium phosphate, $Ca''_3(PO_4)_2$. When the bones are calcined the gelatin burns away and the phosphate remains. The ash is treated with sulphuric acid, which converts the greater part of the calcium into the insoluble calcium sulphate, $CaSO_4$, while impure phosphoric acid remains in solution. This is dried with charcoal powder and intensely heated in a clay retort. The carbon takes oxygen from the acid, forming carbonic oxide gas, CO, which escapes with the hydrogen of the acid. The phosphorus also comes over as gas, but is condensed by passing through cold water.

The changes which occur in the above processes are somewhat more complicated than the description, which is only intended as an approximation, would imply.

Experiments with Phosphorus.

Experiment 1.—Put into a small flask first a quarter of an ounce of ether, then a piece of phosphorus, of the size of a pea. Cork the flask and let it stand some days, frequently shaking it. Decant the liquid; it contains in solution about one grain of phosphorus, and will serve for the following experiments.

Experiment 2.—Pour some drops of this solution upon the hand, and rub them quickly together; the ether will evaporate in a few moments, but the phosphorus will remain upon the hands in a state of minutest division. The more finely it is divided, so much the more easily does it combine with the oxygen of the air. During this combination it diffuses a white smoke and a faint light (it phosphoresces), causing the hands to shine in the dark; hence its name, *phosphorus*, from φῶς, light, and φέρω, to carry. It undergoes slow combustion and is converted into *phosphorous anhydride*, P_2O_3, which rapidly takes up water from the air and becomes phosphorous acid:

$$P_2O_3 + 3H_2O = 2H_3PO_3.$$

Experiment 3.—Moisten a lump of sugar with the solution of phosphorus in ether and throw it into hot water. The surface of the water will glow prettily in a dark room.

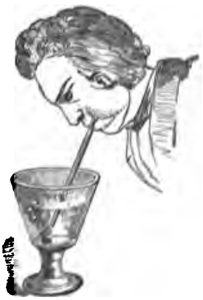
Experiment 4.—Instead of ether use a few drops of carbon disulphide to dissolve the phosphorus. This liquid takes up more phosphorus than ether does, and the solution is highly dangerous. It is the basis of the modern "Greek fire." Its inflammability may be observed in the following way:

Experiment 5.—Pour some of the solution upon fine blotting-paper; the latter *ignites spontaneously* after the liquid has evaporated. The more minutely the phosphorus is divided, so much the more readily it begins to burn.

Experiment 6.—Put a piece of phosphorus of the size of a pea on blotting-paper, and sprinkle over it some soot or pulverised charcoal; it melts after a while, and spontaneously *inflames.* The finely pulverised charcoal causes this combustion, owing to its porosity. It eagerly absorbs oxygen from the air, imparts it again to the phosphorus, and, being a bad conductor of heat, the cooling of the latter is prevented.

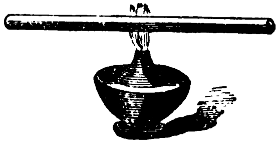
Phosphorus is also *easily ignited by friction*, and is, for this reason, employed in the manufacture of lucifer-matches. The combustible mass is prepared from hot mucilage, at 158° F. (70° C), to which small pieces of phosphorus are added, being thoroughly incorporated with it by constant rubbing till cold. But as the mass, by becoming hard on drying, prevents the access of air to the phosphorus, there must be added some substance rich in oxygen, as black oxide of manganese, nitre, or red-lead, from which the phosphorus can abstract the oxygen necessary for its ignition. $1\frac{3}{4}$ part of phosphorus, 4 of gum Arabic, 4 of water, 2 of nitre, and 2 of red-lead, form a good inflammable mass. A temperature of 149° F. (65° C.)—158° F. (70° C.) is requisite for kindling such matches; in this case the temperature is caused by friction. The coating of the match is thus broken and kindled, and the continued burning is now maintained by the oxygen of the air.

Fig. 61.



Experiment 7.—Put a piece of phosphorus, of the size of a pea, into a wine-glass, and pour hot water upon it until the glass is half filled; the phosphorus melts, but does not ignite, as access of air is prevented by the water. But if air be carefully blown by the mouth through a long glass tube upon the bottom of the wine-glass, *combustion* will ensue, which is visible, especially in the dark.

Fig. 62.



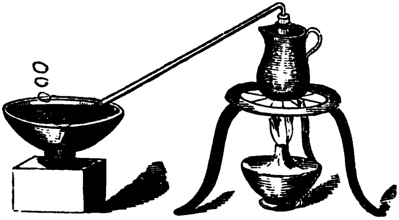
Experiment 8.—Heat gently a piece of phosphorus of the size of a pea, placed in the middle of a glass tube, about twelve inches long. When ignition commences, remove the lamp. While the tube is held horizontally, the combustion is feeble and imperfect, because the heavy smoke, consisting of phosphoric and phosphorous anhydrides, passing off slowly, allows the admission of only a small quantity of air. But the combustion immediately becomes more vivid on inclining the tube, and when the tube is held perpendicularly it is complete, as then the draught of air is most powerful. In this way phosphorus may be oxidized to either degree required; it may be slowly burnt to form phosphorous anhydride, or completely, to form phosphoric anhydride.

Phosphorus, like sulphur, is capable of existing in several different states (*allotropic states*). Besides the ordinary kind the most interesting is that called red, or *amorphous* phosphorus. This is prepared by heating ordinary phosphorus to a temperature of about 240° C. (464° F.) for some time, in a vessel filled with carbonic anhydride to prevent it from burning. Red phosphorus is curiously different in properties from the ordinary kind, being insoluble in carbon disulphide, not poisonous, and only inflammable at a high temperature. It is now used extensively in the manufacture of "safety matches." The tips of these matches do not contain phosphorus, but when rubbed on the outside of the box they take up a little of the red phosphorus with which the side is covered, and combustion ensues.

PHOSPHINE, PH_3.

Experiment 1.—Put into an ounce flask a quarter of an ounce of slaked lime, and a piece of phosphorus the size of a pea, fill it up to the neck with water, and place it in a small vessel containing a strong solution of salt, prepared by adding half an ounce of common salt to an ounce and a half of water. Fit to the flask a bent glass tube, one end of which is made to dip into a basin of water; heat the salt water to boiling, and a gas will be evolved, which, as it issues from the tube and comes in contact with the air, *takes fire spontaneously.* This gas consists chiefly of phosphine, or *phosphuretted hydrogen,* but contains several combinations of phosphorus and hydrogen. If you collect it in a small jar filled with water, it immediately takes fire upon the admission of air. Both the phosphorus and the hydrogen combine with the oxygen of the air, and there results phosphoric anhydride, P_2O_5, and water, H_2O. The first appears as a white smoke, which, when it issues in separate bubbles from the water, rises in rings.

Fig. 63.



Phosphuretted hydrogen, when unburnt, emits the smell of garlic.

In this experiment, the flask is placed in salt water, in order to guard against the ignition of the phosphorus, in case the flask should accidentally break. Salt water, at the strength specified, will not boil under 228° F. (109° C.); consequently, the boiling in the flask is more active than if it had been placed in pure water, the temperature of which, under ordinary pressure, can only be raised to 212° F. (100° C.). The apparatus for heating substances by means of hot water or saline solutions is called a *water* or *saline bath.* By such contrivances extracts are evaporated, and substances dried, which, at a stronger heat, would easily burn, or otherwise be decomposed.

Phosphine, when pure, is inflammable, but *not spontaneously inflammable.* When prepared as above it contains a little of the vapour of another compound of phosphorus and hydrogen

—a liquid—which has the formula P_2H_4. It is the latter compound which is spontaneously inflammable.

COMPOUNDS OF PHOSPHORUS AND OXYGEN.

P_2O_3, Phosphorous anhydride.
P_2O_5, Phosphoric anhydride.

It has been already stated that these oxides can be formed by the direct union of oxygen and phosphorus. The lower oxide and its corresponding acid are unimportant. Both oxides are white, snow-like powders, and they greedily combine with water to form acids.

Phosphoric Anhydride.

Experiment 1.—Ignite a small piece of phosphorus on a plate and cover it with a *dry* bottle or bell-jar. Phosphoric anhydride will be abundantly formed, as a white cloud, which settles on the sides of the bottle and on the plate. Shake the powder into a saucer, and pour a little water on it. A sharp hissing noise will be heard, testifying to the energy with which combination ensues. A solution of phosphoric acid is produced.

Phosphoric Acid, H_3PO_4.

Experiment 2.—Place a piece of phosphorus and a little water in a flask, and carefully heat them until the water boils; then add half an ounce of nitric acid and continue the heat. The phosphorus slowly dissolves, and is oxidized into phosphoric acid. Phosphoric acid prepared in this way is of course contaminated with the excess of nitric acid left undecomposed. This can be removed by evaporating to dryness and boiling the residue with water.

Phosphoric acid is tribasic, containing three replaceable atoms of hydrogen; and inasmuch as one, two, or all three of these atoms may be replaced by metals, its salts are very numerous. With sodium, for instance, the three following salts may be obtained:

NaH_2PO_4, Sodium di-hydrogen phosphate.
Na_2HPO_4, Di-sodium hydrogen phosphate.
Na_3PO_4, Tri-sodium phosphate.

A very great variety of compound salts may be formed by substituting part of the hydrogen with one metal, and part with another. As, for instance, in ammonium magnesium phosphate, $NH_4Mg''PO_4$, in which one atom of hydrogen is replaced by the monad group of atoms, or compound radical, NH_4, and the remaining two atoms by the diad metal magnesium, Mg''.

Experiment 3.—Add ammonia to a solution of sodium phosphate, and then a little solution of magnesium sulphate, $Mg''SO_4$ (Epsom salts). A white precipitate will be produced, consisting of the above-mentioned ammonium magnesium phosphate:

$$MgSO_4 + Na_2HPO_4 + NH_4HO = Na_2SO_4 + NH_4Mg''PO_4 + H_2O.$$

Phosphoric acid or any other soluble phosphate would produce the same effect. Ammonium hydrate and magnesium sulphate therefore serve as a test for phosphoric acid and its salts.

The body of an adult man contains
from 9 to 12 pounds of bones, containing
„ 6 „ 8 pounds of bone ashes, containing
„ 5 „ 7 pounds of calcium phosphate, containing
„ 1 „ $1\frac{3}{4}$ pounds of phosphorus.

Phosphates are also contained in the blood, flesh, and other portions of the body. Whence does it obtain this phosphorus? The answer is, from the meat and vegetables which it consumes. The phosphates occur in bread, in all kinds of grain, in leguminous and many other plants, particularly in their seeds. But how do the plants obtain these salts? By means of the soil. If arable land contained no such salts, no seeds could be produced. If we increase their quantity by mixing ground bones with the soil, we place the latter in a situation to produce a larger quantity of grain; consequently, bones furnish us with a powerful manure.

Besides ordinary phosphoric acid, two other compounds of phosphoric anhydride are known, differing in constitution from ordinary phosphoric acid by the relative quantities of water combined with the anhydride. These compounds are distinguished by the prefixes "pyro" and "meta" from ordinary or *ortho*-phosphoric acid. To form ortho-phosphoric acid three molecules of water are necessary, while for pyro-phosphoric acid two are required, and for meta-phosphoric acid only one:

$$P_2O_5 + 3H_2O = 2H_3PO_4, \text{ Ortho-phosphoric acid.}$$
$$P_2O_5 + 2H_2O = H_4P_2O_7, \text{ Pyro-phosphoric acid.}$$
$$P_2O_5 + H_2O = 2HPO_3, \text{ Meta-phosphoric acid.}$$

These three acids form salts which correspond to them in constitution. The common or ortho-phosphates contain the triad radical PO_4, the pyro-phosphates, the tetrad radical P_2O_7, and the meta-phosphates, which resemble the nitrates in constitution, the monad radical, PO_3.

Experiment 4.—Heat common sodium phosphate to redness in a crucible for some time. Water is driven off, and sodium pyro-phosphate remains:

$$2Na_2HPO_4 = Na_4P_2O_7 + H_2O.$$

Sodium meta-phosphate may be prepared by heating microcosmic salt, $Na(NH_4)HPO_4$ (sodium ammonium hydrogen phosphate). Ammonia and water are expelled, and the meta-phosphate remains:

$$NaNH_4HPO_4 = NaPO_3 + NH_3 + H_2O.$$

These pyro- and meta-phosphates may be dissolved in *cold* water without change, but if the solutions are boiled water is taken up and common phosphates obtained.

When common phosphoric acid is evaporated to dryness, meta-phosphoric, or "glacial" phosphoric acid is obtained. This, when boiled with water, is reconverted to ortho-phosphoric acid:

$$HPO_3 + H_2O = H_3PO_4.$$

Experiment 5.—To solutions of sodium ortho-, pyro-, and meta-phosphates add a little solution of silver nitrate. With the ortho-phosphate, a yellow precipitate, and with the others white precipitates are produced. This test distinguishes the ortho-phosphate from the other two. To distinguish between these, add to each of the salts, first some acetic acid, and then a little white of egg beaten up with water. A white precipitate is produced with the meta-phosphate; none at all with the pyro-phosphate, or ortho-phosphate.

BORON.

B = 11.

Boron is a triad element, but it resembles carbon in its physical characters. Like carbon, it can be obtained as an amorphous powder, and also in the form of octahedral crystals which are nearly as hard as diamonds. Boron is chiefly found combined as boric acid in the steam which issues from fissures in the earth in some parts of Tuscany, and bubbles up through small lakes or basins of water. The acid is obtained by evaporating the water of these lakes by an ingenious arrangement, whereby the heat required for the purpose is supplied by the hot springs themselves.

A sodium salt of boric acid called *tincal* is imported from Thibet; when purified by recrystallization it is sold as *borax*.

Boron is capable of combining directly with many elements, as sulphur, oxygen, and nitrogen.

The oxygen compound of boron is the most important, the others are of little interest.

Boric anhydride, B_2O_3, *and Boric acid*, H_3BO_3, sometimes called *Boracic acid.*

Experiment 1.—Add powdered borax to boiling water until it ceases to be dissolved, pour off the solution from any undissolved borax, and mix it with about half its bulk of strong hydrochloric acid. Flaky crystals are immediately formed, and deposit abundantly as the mixture cools. Collect the crystals in a filter, wash them with a little *cold* water, preserve, and label them boric acid, H_3BO_3.

Experiment 2.—Pour alcohol on some boric acid, or on borax which has been moistened with oil of vitriol. Inflame the spirit, and in either case it will burn with a greenish flame, from the presence of boric acid. Curiously enough, *dry* boric acid may be heated to redness without volatilizing, but if it be boiled with water or alcohol it readily passes off in the vapour of those substances.

Experiment 3.—Heat boric acid in a crucible with a Bunsen's burner. Below a red heat it fuses, and on cooling appears as a colourless glassy mass consisting of boric anhydride, B_2O_3:

$$2H_3BO_3 - 3H_2O = B_2O_3.$$